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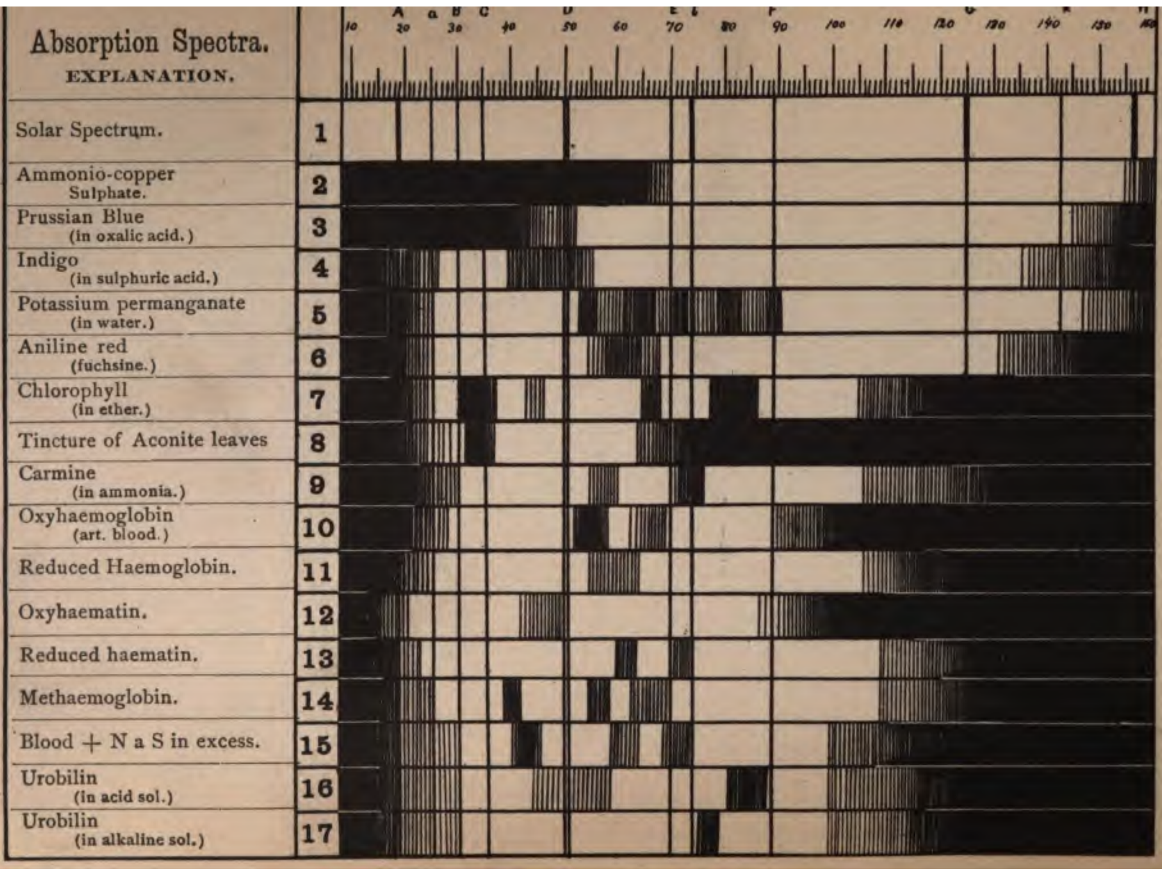

A TEXT-BOOK

OF

MEDICAL CHEMISTRY.

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MEDICAL AND PHARMACEUTICAL STUDENTS
AND PRACTITIONERS.

BY

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PREFACE.

This book is designed especially as a text-book for medical students during their attendance upon lectures. It is believed, however, that it will be found of use to the physician, as a book of ready reference. It is the result of the author's experience as a teacher of Chemistry, during the past twelve years, and its plan, and much of the subject matter, are essentially that given by him in his lectures at the Long Island College Hospital, for the last six years.

While there are numerous good text-books on Chemistry in the market, the medical student often complains that they are either too voluminous for his limited time, and contain a great deal of matter not directly bearing upon the science of medicine, or, on the other hand, are too brief to be of any service to him. It has been the aim, in preparing this book, to avoid these extremes and fill the gap now existing between them.

The plan of the work required a careful selection from the mass of material at hand. The book is largely a compilation, and the author claims little originality in the subject matter it contains, but has used his own judgment in its selection and arrangement.

From the large number of subjects treated, it has been necessary to be very brief in the descriptions of individual substances, but all general principles, and the relations of the facts of the science to medicine have been more fully stated. Much pains has been taken to condense the matter as much as is consistent with clearness.

In Part I are presented such fundamental facts in Chemical Physics, as seemed necessary for the proper understanding of the descriptive parts, and to explain the theory and use of thermometers, the spectroscope, medical batteries, etc.

It has been found best, in the author's past experience, to present the elementary theories of Chemistry before entering upon the natural history of the elements and their compounds;

and this has been done as clearly and concisely as possible in Part II, giving only what was considered necessary to the proper understanding of all the facts to be presented in the later parts. The subjects of Notation, Nomenclature and Chemical Reactions have been dwelt upon as of first importance to the student.

In Part III, the natural history of the Elements and principal compounds are briefly presented, with their physiological and toxicological bearings.

In Part IV, only those organic compounds are treated, with which the physician will be likely to meet. The space that could be given to this part necessitated very short descriptions. The appendix contains some tables and analyses which will greatly enhance its value as a reference book.

The Chemistry of the tissues and secretions have been omitted, as belonging to the domain of physiological chemistry. It would be impossible to mention all the works upon which the author has drawn for his facts.

Besides the credit given in the text, he wishes to acknowledge his indebtedness especially to Cooke's Chemical Philosophy, Barker's, Richter's, Witthaus', Fowne's and Rand's text-books; De Watteville's and Jenkins' books on electricity; Prescott's Proximate Organic Analysis; Wurtz' Dict. de Chem. and Chemie Biologique; Charles' Physiolog. Chem.; Schorlemmer's, Stricker's and Pinner's Organic Chemistry, and the U. S. Pharmacopœia. I would here express my thanks to my friend and assistant, Dr. W. M. Hutchinson, for valuable assistance in preparing the work; also to Mr. W. H. Kent, Ph.D., for similar assistance.

As this work is prepared especially for the medical student, it is to him that I shall trust for its reception, hoping that the labor I have bestowed upon it may help to lessen his.

BROOKLYN, N. Y.,
Sept. 1st, 1885.

CONTENTS.

PART I.

PAGE
Introduction—Definitions 9
  Matter, Properties of 9
  Specific Gravity 11
  Metric System 14
Chemical Physics 16
  Three States of Matter 16
  Constitution of Gases 17
  Laws of Gases 19–21
Heat, Nature of 22
  Boiling Point 24
  Distillation 25
  Latent Heat 26
  Specific Heat 27
  Thermometers 28
Light 29
  Spectrum and Spectrum Analysis 31
Electricity 36
  Theory of Galvanic Cell 38
  Electric Currents 39
  Forms of Cells 40
  Care of Batteries 42
  Ohm's Law 44
  Induced Currents 45
  Induction Coils 46
  Magneto-Electricity 48
  Thermo-Electricity 49
  Static Electricity 50
  Electro-Metallurgy 52
Solution 55
Diffusion of Liquids 56
Dialysis 57
Crystallography 57

PART II.

Theoretical Chemistry 63
  Properties of Molecules 63
  Table of the Chemical Elements 64
  Properties of Atoms 69
  Chemical Notation 71
  Compound Molecules Classified 76
  Nomenclature, Rules of 79
  Nomenclature Simplified 80
  Chemical Reactions and Equations 83
  Stochiometry 86

PART III.

PAGE

Inorganic Chemistry 89
- Classification of the Elements 89
- Hydrogen 91
- Oxygen 94
- Hydrogen and Oxygen—Water 97
- Natural Waters 101
- Water Analysis 104

Chlorine Group 111
- Fluorine 111
- Chlorine 112
- Bromine 116
- Iodine 117

Sulphur Group 119
- Sulphur 119

Nitrogen Group 128
- Nitrogen 128
- The Atmosphere 129
- Disinfectants, Antiseptics and Germicides 132
- Phosphorus 143
- Arsenic 149
- Arsenic Poisoning, etc 153
- Antimony 158
- Bismuth 162
- Boron 164
- Silicon 165
- Carbon 166
- Cyanogen 176

The Alkaline Metals 180
- Lithium 180
- Sodium 181
- Potassium 187
- Rubidium and Cæsium 194
- Ammonium Compounds 194

The Metals of the Alkaline Earths 197
- Calcium 197
- Strontium 201
- Barium 202

The Magnesium Group 203
- Magnesium 203
- Zinc 205
- Cadmium 207

The Iron Group 208
- Chromium 208
- Manganese 210
- Iron 212

The Nickel Group 218
- Nickel 218
- Cobalt 219

The Lead Group 219

PAGE
The Copper Group 224
Copper 224
Mercury 229
Silver 235
Gold 237
The Aluminium Group 238
The Platinum Group 240
The Molybdenum Group 241

PART IV.

Organic Chemistry 243
Homologous and Isologous Series 245
The Paraffines (1st Series) 247
The Olefines (2d Series) 248
Acetylenes (3d Series) 249
Fourth Series 249
Camphors 250
Resins, Gums and Balsams 252
Benzol Series 253
Naphthalene and Anthracene 253
Alcohols 256–259
Ethers 256
Aldehydes 257
Ketones 258
Acids 258
Methyl Alcohol 259
Ethyl Alcohol 261
Alcoholic Beverages 263
Amyl Alcohol (Fusel Oil) 264
Glycerine 265
Hexatomic Alcohols 266
Carbohydrates 267
Cane Sugar 268
Milk Sugar 269
Glucoses 270
Amyloses (starches) 271
Dextrin 272
Glycogen 272
Cellulose 273
Gums 274
Ethers 274
Ether, Ethyl 275
Ethers, Compound 276
Organic Acids 277
Formic Acid 278
Acetic Acid 278
Butyric Acid 280
Valeric Acid 281
Oleic, Palmitic and Stearic Acids 282
Oxalic Acid 283
Succinic Acid 283

Organic Acids—(*Continued*).

PAGE
Malic Acid........ 284
Tartaric Acid........ 285
Citric Acid........ 286
Carbolic or Phenic Acid........ 287
Picric Acid........ 288
Pyrogallic Acid........ 289
Benzoic Acid........ 289
Salicylic Acid........ 290
Hippuric Acid........ 291
Uric Acid........ 291
Organic Bodies Containing Nitrogen........ 291
Amines, Amides, Imides and Nitriles........ 291
Natural Organic Bases, or Alkaloids........ 294
Properties of Alkaloids. Table of Solubilities........ 295
Description of the Principal Alkaloids........ 297
Ptomaines, or Cadaveric Alkaloids........ 300
Glucosides........ 301
Natural Fats and Oils........ 304
Proteids........ 305
Albumins........ 306
Globulins........ 306
Hæmoglobins........ 308
Derived Albumins........ 309
Peptones........ 310
Collagens........ 311
Biliary Compounds........ 312
Vegetable Coloring Matters........ 315
Unclassified Compounds of Animal Origin........ 317
Soluble or Albuminoid Ferments........ 318
Organized Ferments........ 321
Poisons and their Antidotes........ 322

APPENDIX.

Analysis of Blood........ 329
Lymph........ 329
Chyle........ 330
Saliva........ 330
Gastric Juice........ 330
Pancreatic Juice........ 331
Post-mortem Bile........ 331
Human and Cow's Milk........ 331
Tears........ 332
Sweat........ 332
Composition of the Urine of 24 hours, with Pathological Variations... 332
Table of Weights and Measures........ 334
Equivalent Measures........ 335
Specific Gravities named in the U. S. Pharmacopœia, 1880... 336
Solubility of Chemicals in Glycerin........ 337
Solubility of Chemicals in Water and Alcohol........ 338
Incompatibles........ 343
Glossary of Unusual Names of Chemical Substances........ 345

MEDICAL CHEMISTRY.

PART I.

INTRODUCTION.

1. **DEFINITIONS.—Science** is a systematic and orderly arrangement of knowledge. It is founded upon observation and experiment. A **theory** is a deduction from established facts, and occupies a prominent place in science. A **hypothesis** is a supposition brought forward to explain facts or phenomena. **Natural science** treats of the external appearance and internal structure of natural objects. Examples: Botany, zoölogy, and mineralogy deal with the classification and structure of plants, animals and minerals respectively, and are, therefore, natural sciences.

Physical science treats of the properties and phenomena of the matter of which these bodies are composed. **Matter** is anything whose existence is revealed to us by our senses. The **properties** of a body are the peculiar qualities by which it makes itself known to us; as color, solidity, odor, taste, etc. **Physics** is that branch of physical science which treats of the phenomena presented to us by **bodies** or **masses** of matter as such.

Chemistry is that branch of physical science which treats of the ultimate composition of bodies, and the changes which this composition may undergo.

Physics teaches us that water is hot or cold; that it may exist as steam, liquid water, or solid ice. Chemistry tells us that it is composed of two gases called hydrogen and oxygen, in the proportion of two of the former to one of the latter, by volume, and one of the former to eight of the latter, by weight. It also teaches us how we can prove this to be true.

Matter exhibits certain properties which are common to all bodies, and which are hence called the **general properties** of matter; such as indestructibility, impenetrability, divisibility, mobility, gravitation, molecular attraction, chemism and inertia. **Specific properties** are such as are observed in certain bodies only; such as color, hardness, fluidity, transparency, etc.

2. **Matter is Indestructible.**—What we generally term destruction is merely a change of form. When wood burns, it is only changed into invisible gases which go off into the air. These gases may be collected, and by analysis we may get back the elements of the wood again. Whatever changes we may produce in a body or mass of matter, we can neither create nor destroy one particle of it; the same weight of matter remains after the change as before.

3. **Matter is Impenetrable.**—That is, two portions of matter cannot occupy the same space at the same time. Strictly speaking, this only applies to the ultimate particles of a body. In many phenomena bodies do seem to penetrate each other. For instance, the volume of a mixture of sulphuric acid and water, or of alcohol and water, is less than the sum of the volumes before they are mixed. In these cases the penetration is only apparent. The penetration of the water by the alcohol is due to the fact that water, like all other bodies, is porous; that is, it has spaces or interstices between the ultimate particles of which the mass is composed, which are unoccupied by matter, and into which the particles of alcohol crowd themselves.

4. **Divisibility.** Three divisions of matter are recognized, viz: **masses, molecules,** and **atoms.**

A **mass** or **body** is any distinct portion of matter appreciable to the senses.

A **molecule** (a little mass) is the smallest particle of matter that can be obtained by subdividing a mass by mechanical or physical means. It may also be defined as the smallest particle of matter that can exist in the free or uncombined state. A molecule of chalk is the smallest particle of chalk that can exist. The molecules are too small to be seen by the aid of the most powerful microscope, nevertheless their size is approximately known.*

An **atom** is the smallest particle into which any given kind of matter can be divided. The atom is, as yet, a hypothetical body. It is one of the component parts of a molecule, and the smallest particle that can enter into the formation of a molecule.

The atom is, therefore, supposed to be an indivisible solid body, with a definite and unchangeable weight, possessing a definite quantity of force of attraction for other atoms, which force is neutralized by the approach of the requisite number of atoms.

Atoms, when left to themselves, will not, as a rule, remain uncombined, but collect together in groups by their attraction for one another. A collection of atoms forms a molecule, and a collection of molecules forms a mass.

---

* See Art. 15.

5. **Mobility.**—All matter is in a state of constant motion. The motion of masses, or **mechanical motion**, is treated of in works on mechanics. The motions of molecules give rise to the phenomena grouped under the name of the so-called **Physical forces. Light, heat, electricity,** and **magnetism** are different manifestations of the motions of the molecules composing the body which exhibits them. Of the motions of atoms little is known with certainty.

6. **Gravitation** is an attraction which exists between masses of matter. The law of gravitation states that **the force of gravitation is directly proportional to the mass, and inversely proportional to the square of the distance.**

By **mass,** here, we mean the **weight** of matter, and not the **volume.**

By **Specific Gravity** (sp. gr.) is meant the relative weight of equal volumes of bodies, assumed to be under like conditions of temperature and pressure. For the purposes of comparing the weights of equal volumes of different bodies, they are all referred to an assumed standard. The standard for liquids and sòlids is pure water weighed at a temperature of 4° C. (39° F.), the temperature at which it possesses the greatest density or specific gravity.

Density is sometimes used as synonymous with specific gravity. Sulphur, whose specific gravity is 2, weighs twice as much, volume for volume, as pure water; while alcohol, whose specific gravity is 0.825, weighs 0.825 times as much as pure water, volume for volume.

In taking the specific gravity of gases or vapors the standard of comparison is pure, dry air. In chemistry, however, it is more convenient to refer specific gravities of gases to hydrogen gas, and designate this as the **density** and not the specific gravity, which term refers to air as the standard. The density of hydrogen gas is 1.

The density of pure air is 14.42, *i. e.*, air weighs, volume for volume, 14.42 times as much as hydrogen, both gases being weighed at the same temperature and under the same pressure. Density is a very important factor in the study of chemistry, and should be well understood. The sp. gr. of solids may be determined by first weighing the body in the air, and then in water at the required temperature. **A body which sinks in water displaces a volume of water equal to its own, and loses a weight just equal to the weight of water displaced.** The loss in the weight of the body when weighed in water, will, therefore, be the weight of its own volume of water. By dividing the weight of the body in air by the loss of weight in water, the sp. gr. is obtained.

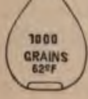
Thus, suppose a body weigh 6.200 grms. in air, and 3.100 grms. in water. The loss of weight in water is 6.200 — 3.100 = 3.100 grms., which represents the weight of the water displaced by the body, or the weight of its volume of water. The specific gravity of the body will then be found by dividing its weight in air, or 6.200 grms., by the weight of an equal volume of water, or 3.100 grms., giving a specific gravity of 2.

When the body in question will not sink, we may attach a sinker to it, whose weight in air and loss of weight when suspended in water are known. The weight of the body in air is taken, the sinker attached, and both lowered into the water and again weighed. The loss in weight will represent the loss of weight of both the solid and sinker. Deduct the loss in the sinker, and the remainder will represent the weight of a volume of water equal to that of the body in question, whence the specific gravity may be found as above. Bodies which are soluble in water may be suspended in some liquid of known specific gravity, in which they are insoluble.

Thus, let it be desired to obtain the specific gravity of a lump of cane sugar. Suppose it weighs in air 100 grms., and in oil of turpentine (sp. gr. .87) 45.62 grms. Loss = 100 — 45.62 = 54.38 grms. 100 ÷ 54.38 = 1.84 as the sp gr. referred to turpentine. Multiply this result by .87, the sp. gr. of the turpentine, and we have 1.6 as the true sp. gr. of the sugar.

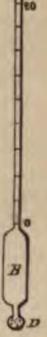
The sp. gr. of a powder is obtained by partly filling a small flask or bottle with it, and weighing both; the weight of the flask empty, as well as the weight of water it will contain, having been previously ascertained. The flask is then filled up with pure water and again weighed. The difference between the last weight and the first will be the weight of the water in the flask. The difference between this weight and the weight of the amount of water the flask will contain when full, will give the weight of the water displaced by the powder, from which the sp. gr. may be obtained as above.

FIG. 1.



The specific gravity of liquids is obtained by means of the flask above mentioned, called a **picnometer** or **specific gravity flask**, made to contain a certain number of grammes or grains of water at a temperature of 60° F., its capacity being marked upon it. (See Fig. 1.) To take the sp. gr. of a liquid, it is only necessary to weigh the flask, filled with the liquid in question brought to the requisite temperature, deduct the weight of the flask, and divide this result by the marked contents of the flask. The specific gravity of liquids is frequently determined also by an instrument called a hydrometer.

FIG. 2.



Hydrometers are long narrow glass or metal tubes, with a bulb near the bottom filled with air, and a smaller one below this containing enough mercury or small shot to make it float upright and to sink it to a convenient depth in water. The hydrometer (See Fig. 2) acts upon the principle of Archimedes, that a body specifically lighter than a liquid sinks in it until it displaces a bulk of liquid whose weight is equal to its own, when it becomes stationary. The long narrow stem composing the upper end of the instrument bears a scale indicating the specific gravity by the depth to which the scale sinks in the liquid.

Hydrometers are of two kinds: (1) for fluids heavier than water, and (2) for fluids lighter than water. Special hydrometers are constructed for use in certain special liquids, and some with an arbitrary scale giving degrees and not specific gravity.

The Twaddell hydrometer, for example, used for liquids heavier than water, is so graduated that the degrees on the scale multiplied by 5 and added to 1000 give the specific gravity compared with water. The **Lactometer** is a hydrometer with a scale specially constructed for the examination of milk. The **Urinometer** is a hydrometer whose scale is made to include the variations in specific gravity found in urine.

To determine the specific gravity of liquids with the hydrometer, it is only necessary to drop the instrument into the liquid, which must be at the temperature for which the instrument is constructed, which, in this country, is usually 60° F., and then read off the specific gravity on the scale at the surface of the liquid. (See Appendix, for table of sp. gr. of chief liquids of U. S. P.)

7. **Molecular Attraction.**—Molecules attract one another, as well as masses. When molecules of the same kind attract one another, they form a homogeneous mass, and the force acting between them is called **cohesion**; when the molecules are unlike it is called **adhesion.** A body on being thrust into water comes out wet, because the water adheres to the body; but if you attempt to pull it apart, cohesion keeps it together.

8. **Chemism.**—Atoms attract one another by a force called

chemism or **chemical affinity**. Chemism holds atoms together to form molecules. It is to the molecule what cohesion is to the mass. Like cohesion it only acts across inappreciable spaces.

9. **Inertia** is that property of matter by virtue of which it has no power in itself to change its condition. [This term is often incorrectly limited to the tendency of a body when in motion to remain in motion, and when at rest to remain so.] Chemical or physical changes never take place without the intervention of some external agency.

10. **Extension** is that property of matter by virtue of which it occupies space. Its relative degree is obtained by means of weights and measures. The system of measures and weights most in use in all scientific works is that known as the **metric** or **decimal system.** In most American and English medical books, though not in all, the English system is employed, while all nations of continental Europe use the metric system; so that it is requisite that the student should be familiar with both. The unit of the metric system is the metre. The **metre** is the length of a platinum bar deposited in the public archives of France, and is 39.37 inches in length. This measure was obtained by taking the 10,000,000th part of the quadrant of a meridian of the earth, or of the distance from the equator to the pole. The ratio of increase and decrease of the system is decimal, and this system is consequently sometimes known as the **decimal system.** The multiples in all the measures and weights are denoted by Greek numerals, used as prefixes, as will be seen by reference to the tables on page 2 of the Appendix. Thus, in measures of length we have the metre, **deka**metre, **hecto**metre and **kilo**metre. The subdivisions are denoted by the Latin numerals, thus: metre, **deci**metre, **centi**metre and **milli**metre.

Greek.	*Latin.*	*English.*
Deka.	Decem.	Ten.
Hekaton.	Centum.	One Hundred.
Chilios.	Mille.	One Thousand.
Murias.		Ten Thousand.

The decimal subdivisions and multiples give a simplicity to the tables which enables them to be easily learned and remembered, and brings into use in all calculations the easiest arithmetical processes. In square measure, measures of capacity, and measures of volume, but few denominations are in common use, and these only are given in the table.

The **litre**, or cubic decimetre, is equal to 1000 cubic centimetres. The cu. centimetre (abbreviated c. c.) and tenths of c. c. are the only denominations in common use for quantities less than the litre. The half-litre and quarter-litre are also used. Measures of weight are derived from measures of volume, as follows:—

The **gramme**, the unit of weight, is the weight of 1 cubic centimetre of distilled water, weighed in a vacuum, and at the temperature of 4° C. (39.2° F.), the temperature at which water has the greatest density.

Theoretically, the unit of weight is derived from that of capacity, as above; but the gramme is really determined by reference to an original standard kilogramme weight adopted by the French government.

The capacity of vessels is determined, not by measure, but by weighing the water they will hold at a selected temperature.

A litre measure is, therefore, a vessel that will hold exactly a kilogramme (1000 grms.) of pure water at the temperature at which it is to be used. Measuring instruments are usually made of glass, and are made to hold their marked contents at a temperature of 60° F. (15.5° C.)

The following table will be found convenient to memorize, in order to facilitate mental calculations where approximate values only are desired:—

1 Metre = 39.37 in. = 3¼ feet.
1 Decimetre = 10 centimetres = 4 in.
1 Litre = .88 quarts = 1¾ pints.
1 Gramme = 15½ grains.
1 Kilogramme = 2 lbs. avoirdupois.
1 Grain = .065 grammes.
ʒj or fʒj = 4 grms. or 4 c.c.
℥j or f℥j = 30 grms. or 30 c.c.

11. The Metric System in Prescriptions.—In prescription writing it will be found convenient to adopt the following rule where the doses have been learned in grains: Make f℥j equal 30 c.c. *In a two ounce mixture, when a teaspoonful is to be given as a dose, write as many grammes as there are grains or minims required in each dose.* Thus, suppose it be desired to write for a two ounce mixture containing 15 grains of potassium bromide in each fʒj, it would be written:—

Grammes.
℞. Potass. Bromidi...........15|000
Aquæ.......................60|000 or 60 c.c.
M. Sig.—4 c.c. ter die.

A f℥ iij mixture would be written with one and one-half times as many grammes as there are grains required in each dose:—

Grammes.
℞. Potass. Bromidi....................22|500
Aquæ....................................90|000 or 90 c.c.
M. Sig.—4 c.c. ter die.

The perpendicular line is used here to occupy the position of the decimal point, and separate grammes from milligrammes.

It is customary among some pharmacists to weigh all quantities written as above, instead of measuring them. In such cases, it is well to remember that **liquors** are very nearly the same specific gravity as water, excepting those containing the salts of the heavy metals. The **syrups and glycerita** have a sp. gr. of about 1⅓. The **tinctures** have a sp. gr. of about .9. In cases of doubt about the sp. gr., direct the quantity in c.c., when it will be measured.

CHEMICAL PHYSICS.

12. We have already spoken of the divisions of matter into masses, molecules and atoms. The physicist deals principally with the properties of bodies or masses of matter which depend upon behavior of the molecules, without inquiring of what atoms they are made up. Hence, when a piece of iron is moulded into this or that form, or melted and cast, or filed, or ground to powder, it still remains iron. It may be hot or cold, it may be charged with magnetism or electricity, but it is iron still, and nothing else. These properties of color, solidity, tenacity, malleability, fusibility, power of being magnetized, and of conducting electricity, all depend upon the behavior of the molecules of iron, and are called **physical properties.**

13. **Three States of Matter.**—The same body may exist in three states; as a solid, liquid, or gas. Most solid bodies can be changed into the liquid or gaseous state by the application of heat to them. For example, when ice is heated, it melts and assumes the liquid form; if we apply heat to the water, it expands, is converted into steam, and tends to escape from the vessel containing it. If we confine the steam, it exerts a strong pressure upon the walls of the vessel, which increases as the temperature increases. This and other similar facts have led to the adoption of the word **force** to express the nature of heat. The term **heat force** may often be found in works on physics. The capacity which this force possesses of performing work is known as energy. Cohesion has been defined as that force or attraction

which acts between similar molecules to hold them together to form a body or mass. From the above illustration, we see that heat, whatever it may be, acts in opposition to cohesion, and even destroys it altogether, driving the molecules apart, and finally off into space. We have in this, then, the explanation of the three states of matter. In the **solid state,** cohesion is greater than the opposite repellant force of heat, and the molecules are in very close apposition. In the **liquid state,** cohesion and the repellant force are nearly equal, the molecules being free to move in any direction, and obey the law of gravity; *i. e.*, they all tend to seek the lowest level. In the **gaseous state,** the repellant force of heat has entirely overcome the attraction of the molecules for one another, hence, they tend to fly off into space, and will do so unless confined in a vessel. If we withdraw the heat from a gas, or if we compress it sufficiently, we may reduce it to a liquid, and in many cases to a solid; *i. e.*, we bring cohesion into play again.

It will be easy to understand from the above, that the molecules must be farther apart in steam than in water, and in water than in ice. What is true of water, is true of all bodies capable of existing in the three states; for heat expands all bodies, whether solid, liquid, or gaseous, although not to the same degree. Gases expand more than liquids or solids for the same change of temperature. Gases are more simple in their constitution than either of the other states, and hence we get a better knowledge of their structure.

14. **Constitution of Gases.—Law of Avogadro or Ampére.**—This law states that **equal volumes of all bodies in the state of gas, and at the same temperature and pressure, contain the same number of molecules.**

As a necessary deduction from the law, it follows: 1st. That **all gaseous molecules occupy the same space.** 2d. That the **weights of equal volumes of any two gases are to each other as the weights of their molecules**; or, in other words, **the specific gravities of any two gases must be to each other as the weights of their molecules.** This law is of vast importance to the chemist, and is considered the basis of most of the modern notions of chemistry. Although it would be out of place here to enter into such a discussion, this law is capable of mathematical proof, starting with the assumption that masses are composed of molecules in a state of motion.

15. **Size and Weight of Molecules.**—That molecules actually exist, and that they are in constant motion in straight

lines within the gas, there seems to be at present little room for doubt. Various recent experiments, drawn from many sources, have given us proofs of these facts. Starting from certain well established facts, physicists have been able to calculate the absolute number of molecules in a given space, their absolute weight, size, velocity, and the spaces between two neighboring molecules. What was at first a mere hypothesis is fast becoming a demonstrated fact.

According to these calculations, a cubic centimetre of air contains twenty-one trillions of molecules; and according to the law of Avogadro, all other gases must contain the same number in the same volume. 10 trillions of air, or 144 trillions of hydrogen molecules, will weigh 1 milligramme. The mean velocity of the molecule of air at 0° C. (32° F.) is 485 metres (1591 ft.) per second; and of a molecule of hydrogen gas is 1844 metres (6050 ft.) per second. Of course, with this inconceivable number of molecules in the small space of one cubic centimetre, and all moving at the velocity mentioned, no one molecule could move long in one direction without colliding with another molecule. The number of shocks that each molecule receives, in the case of hydrogen gas, has been calculated to be 9480 millions per second, while the mean distance a molecule moves in its path before colliding, is about .0001855 mm., which may be taken as the distance between two molecules. The diameter of the water molecule = .00000044 mm. Free path = .0000649 mm. Although these numbers give us no real conception of the magnitudes they represent, they are given here to show the tendency of research, and the advances being made. These numbers apply to gases only.

16. **Radiant Matter.**—When the pressure is removed from gases, the molecules are allowed to go farther apart; and while the distance between the molecules of air at the ordinary atmospheric pressure is a little less than a ten thousandth part of a mm., it may reach several centimetres when the pressure is reduced by a vacuum pump to a millionth of an atmosphere. Such a rarefied gas has new properties, and has recently received the name of **radiant matter**, by Mr. Crookes. Under the influence of a ray of light, or an electric discharge sent through the rarefied gas (pressure one millionth of an atmosphere), the charged molecules, in the case of electricity, seem to launch forth in a direct line from the negative pole towards the positive. If a light disc be placed in this path, it is bombarded by these molecules on the side presented to the negative pole, and thus it receives a sufficient impulse to render the motion visible to the eye. In this manner the movement of the radiometer is explained.* We are now in a position to understand the laws of gases.

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* Wurtz, Dict. Sup., p. 249.

17. **Law of Mariotte.**—This law is stated as follows: **The volume of a confined gas is inversely proportional to the pressure brought to bear upon it**; that is, the less the pressure the greater the volume, and the greater the pressure the less the volume. Stated algebraically this law is: $P : P' :: V' : V$, whence $V' = \frac{PV}{P'}$, in which expression P and P′ stand for two different pressures, and V′ and V for the corresponding volumes. The pressure exerted upon any mass of gas is usually that of the surrounding atmosphere, and is measured by a barometer. The normal pressure of the air is about 15 ℔s. to the square inch, or such that it will support a column of mercury 760 mm. (30 inches) in height. In the expression of pressures, as measured by this instrument, the height of the column of mercury is employed instead of the weight of the column. Thirty inches of mercury is used, instead of 15 ℔s. pressure. Substituting this form of statement in the above expression, it becomes $V : V' :: H' : H$ and $V' = \frac{VH}{H'}$, or $VH = V'H'$ in which H and H′ stand for the height of the barometric column.

Example.—A certain volume of air, at a pressure of 742 mm., measures 540 c. c.; what will it measure at 760 mm? By the above formula we will have $V' = \frac{540 \times 742}{760} = 527.2$ c. c.; or 540 c. c. : x :: 760 : 742 in which $x = 527.2$.

Since a gas undergoes such great variation in volume under varied pressures, it is necessary to state the pressure when a given volume of gas is mentioned, or as is commonly done, to have an arbitrary standard of pressure under which all gases are supposed to be measured when not otherwise stated. The pressure which has been adopted is 760 mm., or 30 inches of barometric pressure. This is called the standard pressure.

18. **The Law of Charles.**—Not only does the volume of a gas vary with the pressure, but it also varies with the temperature. The hotter the gas the greater is its volume, and the cooler the gas the smaller the volume; or, as this law is usually stated, the volume of a gas varies directly with the absolute temperature. This is known as the law of Charles.

All gases expand or contract equally for the same increase or decrease of temperature. A gas expands $\frac{1}{273}$ of its volume in passing from 0° to 1° C. or $\frac{1}{460}$ of its volume for one degree Fahrenheit.

Now, since a gas expands $\frac{1}{273}$ of its volume at 0° C., for every degree increase of temperature above zero, we may regard a gas at 0° C. as having been warmed through 273° C. In other words, 273° below zero mus

$\sqrt{D'} : \sqrt{D}$. Or the velocities of the molecules are to each other inversely as the square roots of their densities, which corresponds to the statement of Graham's law. The densities of hydrogen and oxygen are known to be 1 and 16 respectively, and, therefore, the velocities of their molecules must be as the square roots of these numbers or as 4 to 1; or the molecule of hydrogen must move 4 times as fast as that of oxygen. The comparative velocities of other molecules may be calculated in the same way, from their densities.

22. **Nature of Heat.**—As nearly as we can conceive, the phenomena and sensations to which we apply the term heat, are the manifestations to our senses of the motions of the molecules of matter, which we have partially discussed in the last sections. Besides producing the sensation of heat, it acts variously on bodies; it boils water, melts iron, makes the metals give out light, electricity, etc.

It was formerly supposed that heat was a form of matter—a subtle fluid which could flow from one part of a body to another, or through the air. We still retain some of the forms of expression used while that theory was held; such as **conduction, convection, absorption, emission, radiation,** etc., the theory itself being entirely abandoned. We now regard heat as a manifestation of one form of molecular motion. The more rapid the motion of the molecules of any given body, the higher will be the temperature.

Many of the phenomena of heat may be beautifully illustrated by means of a ball attached to an elastic India rubber string, held in the hand. The ball may represent the molecule, the string the elastic cohesive force, which acts between molecules, and the force applied by the hand to make the ball revolve about it, the force of heat.

23. **The Sources of Heat** are the sun, stars, interior of the earth, chemical action, and the conversion of mechanical motion or molecular motion into heat. The earth receives but about the two thousand millionths of the heat of the sun, and the fixed stars are estimated to furnish about $\frac{4}{5}$ as much heat as the sun.* The amount of heat annually received from the sun would melt a layer of ice surrounding the earth, 101 feet thick.* The internal heat of the earth has some effect upon our temperature, but not a very important one, as the crust of the earth does not conduct heat well. Chemical action is the most important artificial source of heart. When the heat thus produced is intense, and accompanied by light, the bodies are said to burn. In an ordinary fire, the heat produced is due to the chemical action

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* Rand, p. 34.

going on between the oxygen of the air, and the carbon of the fuel. Animal heat is largely due to a similar cause; that is, to chemical action going on in the muscles, glands, brain, and, in fact, all the tissues; not, as formerly taught by Liebig, by direct combustion of the food by the oxygen of the air, but by oxidation of the different tissues whenever they are called upon to exercise their functions.

Mechanical force may be converted into heat; as when the axle of a railroad coach becomes hot; or when the brakes are applied to the wheels; or when a piece of steel strikes a piece of flint. These phenomena may be imitated with the ball and string above mentioned, by allowing some one to strike the ball while it is swinging, so as to drive it faster on its course. This is what takes place when the steel comes in contact with the flint.

24. **Mechanical Equivalent of Heat.**—There is an intimate relationship between heat and mechanical motion. They are capable of being converted, the one into the other. The friction of the match against a roughened surface produces enough heat to ignite it. Heat, on the other hand, is converted into motion by the expansion of steam, which drives the steam engine. It has been determined that a certain amount of heat has its exact equivalent of work. The unit of work is the foot-pound—the force required to raise one pound one foot high; and the unit of heat is the heat necessary to raise one pound of water from 0° to 1° C. This amount of heat, if it could all be made to do mechanical work, would be sufficient to raise 1390 lbs. one foot high; or the thermal unit is equivalent to 1390 foot-pounds. If we adopt as the thermal unit, the heat necessary to raise one pound of water from 32° to 33° F., it is equal to 772 foot-pounds.

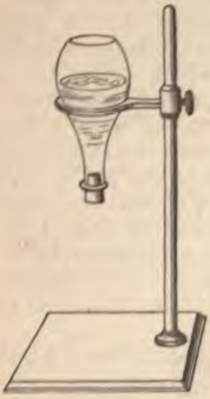
25. **Effects of Heat.**—One of the first effects of heating a body is to expand it. This is the effect of heat upon all bodies, except in a few cases of apparent exception. Silver iodide is a notable exception to this rule. To understand just what takes place when heat expands a body, let us return to our ball and string; the more force we apply to it from the arm, the more rapidly the ball will move; as it does so, the more it stretches the string and the larger the arc it describes. If we apply enough force the string will break, and the ball will fly off into space.

26. **Melting and Freezing Points.**—When heat is applied to the molecules, it drives them faster through their paths and further apart from one another, until finally the cohesive force is

stretched to its utmost, and is on the point of breaking; then the body melts and becomes a **liquid**, in which state the two forces are nearly equal, with cohesion slightly predominating. The temperature at which a body, usually solid, passes into a liquid state is called its melting point. If heat is abstracted from a body that ordinarily exists in a liquid state, cohesion more and more overcomes the heat force between the molecules until the body contracts and passes into the solid state. The temperature at which this takes place we call the freezing point of the body. If more heat is now applied, the cohesive force (represented by the elastic string above) gives way, the molecules begin to fly off into space, and we say the liquid boils or passes into the **gaseous state.**

27. **Boiling Point of Liquids.**—The temperature at which a liquid gives off vapor rapidly from the whole liquid is constant for that liquid, and is called its **boiling point.** In giving a description of a liquid the boiling point is usually given. The boiling point of water is 100° C. (212° F.), and is nearly constant when this occurs in an iron vessel in the open air. It is slightly higher in a glass or other vessel with polished walls, because the steam adheres to such a surface until it becomes a little above that point, or slightly superheated. The pressure exerted upon the surface of the liquid may vary the temperature of the boiling point, by resisting the expansive force of the molecules and aiding cohesion to keep them together. When considerable pressure is brought to bear upon a boiling liquid, its temperature rises until the tension of the steam overcomes this pressure, and it becomes **superheated.** Conversely, if the pressure be removed, or lessened, the temperature at which the liquid will boil is lowered. This may be illustrated by the experiment known as the **culinary paradox,** as follows:—

FIG. 4.



Take a flask of water, and after boiling it vigorously for a few minutes, so that the rising steam may expel most of the air, cork it up with a tightly-fitting cork and the boiling will cease; remove the lamp and turn the flask bottom upward. (See Fig. 4.) By pouring cold water on the flask the steam is condensed, the pressure is removed from the water, and it boils vigorously. Now, allow the steam to fill the flask above the water, and it ceases boiling again. A dash of cold water will again make it boil. This may be repeated until the water becomes cool enough to be held in the hand with comfort.

The boiling point of water varies, then, with the height of the barometer, or with the pressure of the atmosphere upon it. As the height of the barometer falls in ascending from the surface of the earth, the boiling point of water must fall also. An ascent of about 1080 feet lowers the boiling point of water 1° C. By this means the height of mountains may be determined within a few feet. When a liquid passes off into vapor, it is said to **evaporate.** Evaporation takes place slowly and imperceptibly from liquids at temperatures below their boiling points. In the case of water, for example, there is some evaporation even from the surface of snow and ice. The moisture of the air, from which clouds, rain, snow, dew, etc., are formed, is carried up by this imperceptible evaporation.

28. **Distillation and Sublimation.**—When a liquid is rapidly evaporated and condensed again by means of cold, the process is called **distillation.** When water containing solid matters in solution is evaporated, the solids remain in the vessel while the water only is given off. By means of this fact, we are able to prepare perfectly pure water by distillation. When a mixture of two or more liquids is heated, that having the lowest boiling point begins to evaporate or distill first, leaving the others behind. During the rapid evaporation of a liquid, its temperature remains constant at the boiling point until it is all evaporated. To separate a mixture of liquids having different boiling points, it is only necessary to heat the mixture until that having the lowest boiling point begins to boil, and allow it to remain at that temperature until it has all passed over and been condensed. The condenser is now removed and another attached; the temperature is now raised until another portion of the mixture begins to distill over, and so on until the liquids are separated. The first process seldom effects a perfect separation, owing to some heavier liquids being carried over by the lighter ones; a second or even a third distillation is often necessary. The above process is called **fractional distillation.**

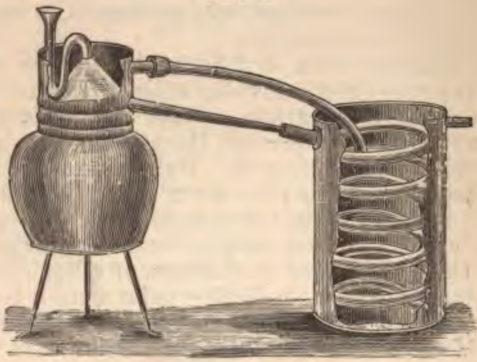
By the term **destructive distillation** is meant, a distillation usually of dry substances, so as to destroy them and obtain liquids or gases; as, for example, the distillation of coal for the purpose of preparing illuminating gas, and liquids to be used for various purposes; and the distillation of wood to prepare vinegar, wood spirit, etc. Distillation is carried on in a retort or still, and the vapor is condensed in a worm or condenser. The retort, or still, is the vessel in which the liquid is heated, and is made of glass, copper, iron, or platinum. (See Fig. 5.)

The heat is applied to the retort or still until the liquid boils. The vapor from the boiling liquid passes through the beak of the retort into the condenser, which is always kept cool by means of cold water.

A few solid bodies when heated do not melt and form liquids, but pass directly into the state of vapor. Such bodies are said to **sublime**, and the process is called **sublimation.**

Iodine, sulphur, camphor, and ammonium chloride, are examples of bodies which may be sublimed, and this process is usually employed for their purification.

FIG. 5.



29. **Latent Heat.**—It is evident that a part of the heat force applied to a body is used up in overcoming the force of cohesion and in expanding the body, and does not appear in the actual moving power of the molecules. In our ball and string illustration (Art. 22), a part of the force applied to the string by the hand is expended in stretching the string, or finally in breaking it, and does not appear in the moving power of the ball. This force which is expended in overcoming cohesion and in keeping the molecules apart, does not appear in the temperature of a body. When air is heated and allowed to expand, about two-sevenths of the heat force is used up in expanding it. If we apply heat to a vessel containing ice, the temperature of the water formed is the same as that of the ice, although a considerable heat has been absorbed in the melting process.

When water is boiled it does not rise above 100° C. (212° F.) however hot the fire is; it remains at 100° C. until it is all evaporated. What has become of all the heat applied to the water? It has been used to drive the molecules farther apart, and is locked up in the water in the first case, and in the steam in the second. The heat thus expended is called **latent heat. It requires heat to convert a solid into a liquid, or a liquid into a gas.** The reverse is equally true; that **heat is given out when a gas becomes a liquid, or a liquid becomes a solid.** Freezing mixtures are based upon this law. A mixture of salt and ice freezes, because the salt melts the ice by its affinity for water; but heat is absorbed by a solid when it becomes a liquid, and is taken up in this case from whatever happens to be in contact with the freezing mixture or vessel containing it. Ice machines operate on the principle that when a liquid evaporates, it absorbs large quantities of heat. A very volatile liquid is made to evaporate rapidly by removing the pressure from it by pumping out the vapor as fast as formed. A very low temperature is produced in this way in and about the vessel containing the liquid. Place a few drops of ether upon your hand and allow it to evaporate, and you will have a very good illustration of the heat absorbed by an evaporating liquid. Mitchell obtained a temperature of —146° F. with ether and solid carbon dioxide. Vatterer produced a temperature of —220° F. with liquid nitrous oxide and carbon disulphide.

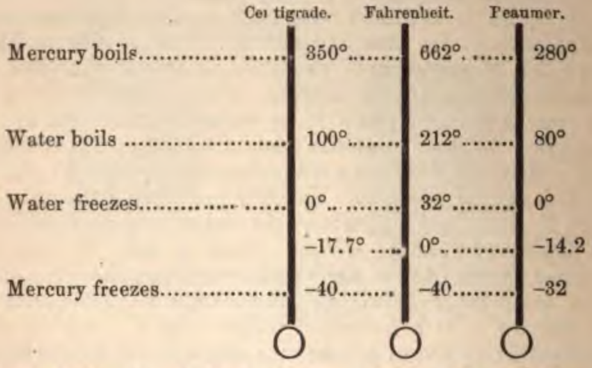
30. **Specific Heat.**—When equal weights of two given bodies are exposed to the same source of heat, they do not both rise in temperature with the same rapidity. That is, it takes more heat to raise a pound of water from 0° to 1°C., than it does to raise one pound of mercury through the same change of temperature. The relative amount of heat required to raise equal weights of substances through equal degrees of temperature, is called their **specific heat.** Water has the highest specific heat of any known substance except hydrogen gas. The unit of specific heat is not everywhere the same. Some use as the unit of heat the heat required to raise 1 pound of water from 32° to 33° F., others from 0° to 1° C. The latter is the one most commonly used, and is called the **thermal unit.** In France, the unit is the **calorie,** which is the heat required to raise 1 kilogramme of water from 0° to 1° C. 1 Calorie = 2.2 thermal units. 1 thermal unit = 0.45 calorie.

31. **Temperature.**—From the preceding article it will be seen that the temperature of a body is entirely distinct from the amount of heat it contains. Temperature may be defined as the

tendency a body possesses of imparting heat to surrounding bodies. It is this tendency which gives to our senses the impression that the body is **hot** or **cold**; it is, therefore, the measure of its **sensible heat,** or heat that is appreciable to our senses. The heat which a body contains is made up of its sensible heat plus its latent heat. The temperature of a body is increased or diminished by adding to or withdrawing from it sensible heat. As the number and weight of the molecules of a given body are constant, the variations of temperature must mean a variation in the velocity of the moving molecules. In our ball and string illustration it will be indicated by the velocity of the ball in its path.

32. **Thermometers.**—Temperature is measured by an instrument called a thermometer. Thermometers are usually

FIG. 6.



constructed of a closed glass tube, provided with a bulb at one end, containing a liquid whose expansion or contraction is used to indicate the temperature. It is also provided with a scale to mark the amount of contraction or expansion taking place in the liquid. Liquids are usually chosen, as solids expand too little and gases too much to be convenient. The liquids commonly used are mercury or alcohol; of these the former is most extensively used, because of the long range of temperature between its freezing and boiling points. Pyrometers are instruments for measuring very high temperatures, and are constructed of metal, or fire clay, whose melting point is very high.

The thermometric scales in common use are the Fahrenheit, Celsius or Centigrade, and Reaumer. The differences in these scales may be seen at a glance by reference to Fig. 6. There are two fixed points in all of them—the temperature of melting ice, and that of the steam from boiling water. These two points must be determined on every instrument by actual trial; these points are marked on the glass with a file or diamond. It then remains only to divide the space between them into a certain number of degrees, according to the scale adopted. In the Centigrade or Celsius and in the Reaumer the freezing point of water is marked 0°, while in the Fahrenheit it is marked 32°.

The point at which the mercury rises in the tube when the latter is plunged into steam from boiling water, is to be marked 100° in the Celsius, 212° in the Fahrenheit, and 80° in the Reaumer. There remains, then, simply to divide the space between these points into 100 equal parts in the first, 180 in the second, and 80 in the third. Chemists generally have adopted the Centigrade scale, although some still adhere to the Fahrenheit, which is the one in common and almost universal use among the laity in this country. The Reaumer is not much used in this country, and we shall not use it in this book. It is not difficult to change the readings from one to the other scale.

It will be seen that 100° C. = 180° F., 1° C. = 1.8° F., or 1° C. = $\frac{9}{5}$° F. and 1° F. = $\frac{5}{9}$° C. We must remember, however, that the 0° mark in the Fahrenheit scale is 32 degrees below that of the Celsius; hence, in converting degrees F. to degrees C. we must first take from the reading 32° and reduce only those above the freezing point of water. While in changing degrees C. to F., we must add 32° to the result, to obtain the true reading.

Thus, 10° C. $= 10 \times \frac{9}{5} = 18 + 32 = 50$: and 41° F. $= 41 - 32 = 9 \times \frac{5}{9} = 5$° C. Or, multiply degrees C. by $\frac{9}{5}$ and add 32 = degrees F.; and multiply degrees F. by $\frac{5}{9}$ after subtracting 32 = degrees C.

The Centigrade and Fahrenheit degrees will both be given in this book.

LIGHT.

33. The second so-called physical force which plays an important part in many chemical phenomena, is light. According to the best conception we have of light, it is the effect upon the optic nerve produced by undulations of an exceedingly subtile and highly elastic form of matter, called the **luminiferous ether.** That this ether really exists, pervading the spaces between the molecules of all bodies, so many million times more elastic than air, so light that it offers no appreciable resistance to the earth moving 1100 miles a minute through it, is a hypothe-

sis, merely, which has been advanced to explain certain well-known facts. That light passes from the sun and stars to the earth, no one can doubt; and yet, without some such assumption, we cannot conceive how it does pass, unless we hold to a former view, which taught that light was in itself a form of matter, without weight, given off by luminous bodies, and which is able to pass through glass, water, rocks, etc. These properties are contrary to all known laws of ordinary matter, as also are those of the ether assumed in the other theory.

It is believed that undulations may originate in certain motions of molecules or atoms; that these motions or undulations are communicated to the ether and conveyed upon it in the form of waves; and that the movement described by a given particle of ether is in the main an oscillation at right angles or perpendicular to the direction of the ray, or beam.

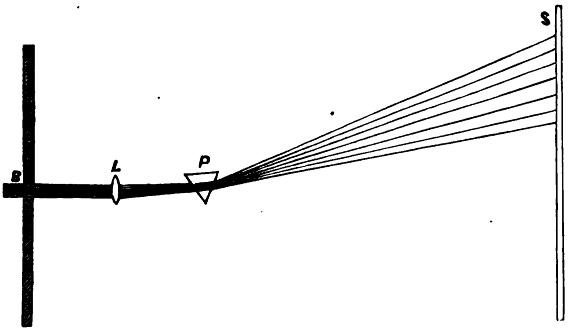
34. **Transmission of Light.**—The motion of a ray of light travels along the line of particles, very much in the same way that it passes along the line of ivory balls placed in contact on the billiard table, when another ball strikes the end one directly in line with the rest. The motion or impulse passes along the line, but the one at the opposite end is the only one seen to move. It may be illustrated by placing a ruler on the table, holding it firmly in place with a marble in contact with one end, and with a hard body striking a short quick blow at the other. The stick does not move, as a whole, but the jar will be felt to pass under the hand, and the marble will move from its place.

35. **Color and Intensity.**—It is evident that these oscillations may differ in rapidity or in their amplitude; *i. e.*, comparing it to the movements of a pendulum, it may vibrate rapidly or slowly, and it may swing a long distance or a short one. Upon the extent and rapidity of these oscillations depend two important differences in the effects of light on the organs of vision, viz., color and brilliancy; the brilliancy depending upon the force of the blows upon the retina, and the color upon the number in a given time. We have an analogous fact in sound. Here we can more easily demonstrate the truth of the fact that the intensity of the sound depends upon the amplitude of the vibrations of the molecules, while the pitch depends upon the number of waves or pulsations which reach the ear in a given time. From well established data, we are able to calculate the rapidity of the oscillations which produce the different sensations of color, and the corresponding lengths of the ether waves. Some of these results are expressed in the following table:—

Color.	Length of Waves in Fractions of a Millimetre.	Number of Oscillations in one second
Red.	650 millionths.	477,000,000,000,000
Orange.	609 "	506,000,000,000,000
Yellow.	576 "	535,000,000,000,000
Green.	536 "	577,000,000,000,000
Blue.	498 "	622,000,000,000,000
Indigo.	470 "	658,000,000,000,000
Violet.	442 "	699,000,000,000,000

The color of an object depends upon the character of light it reflects or transmits to the eye. A beam of white light is composed of a variety of different colored lights mingled together, as can be shown by passing it through a prism of glass having an angle of 60°, by which it is decomposed into its component colors. When a body looks red to us, it is because it absorbs or destroys all the oscillations of the white light except those which give us the sensation of red light. If it be blue, only the vibrations that give us the sensation of blue light are reflected. Some bodies and solutions reflect one color, and transmit another very different one. The color transmitted is usually the complement of the one reflected; *i. e.*, if the two lights are mixed together they produce the sensation of white light. If a solution of nickel salt and one of cobalt are mixed together cautiously, the color of the one mixes with that of the other so as to form a colorless solution, because the colors are complementary.

FIG. 7.



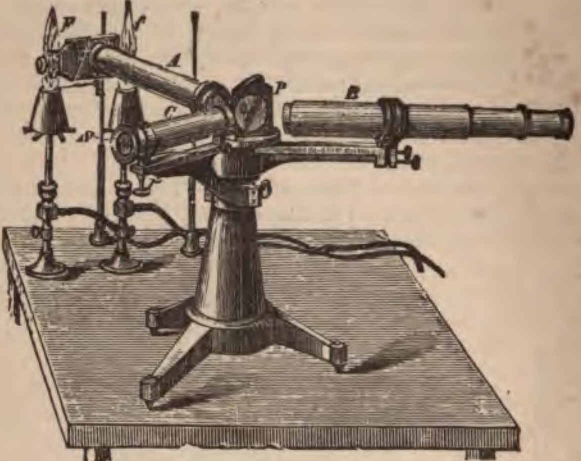
36. **The Spectrum.**—When a beam of white light is passed through a prism, as represented in Fig. 7, it is not only refracted,

that is, bent from its original course, but the colors of which it is composed being unequally bent, are separated one from another. If now we allow them to fall upon a white screen, S, Fig. 7, they produce a series of blending tints upon it, which are called a **spectrum.**

The red rays, which are least bent from their course, are said to be the **least refrangible,** while the violet are the **most refrangible.** Intermediate between these colors we find the orange, yellow, green, blue, and indigo. The prism thus gives us an easy means of analyzing a beam of light, to show the character of the rays producing it. Such observations are usually conducted by means of an instrument called a spectroscope.

37. **The Spectroscope.**—Figure 8 will illustrate the construction of the spectroscope. The light is received from the source of light (F), through a very narrow slit, regulated by a screw; it passes through the tube (A), called the collimator, and is directed upon the prism (P) which may be made of flint glass, or it may be made hollow and filled with bisulphide of carbon. In some instruments there are several flint glass prisms so arranged that the light is made to pass through all of them, so as to secure a wider dispersion of the rays than can be obtained with one prism.

FIG. 8.



The beam of light, after traversing the prisms, is viewed with the telescope (B). For purposes of comparison, an additional tube (C) is attached, by which another light may be thrown upon the surface of the prism so as to be reflected through the telescope by the side of the light from F.* Many instruments now made are so arranged that the ray of light is not bent from its course in passing through the prism, and these are called **direct vision spectroscopes.**

38. **Bright Lines.**—When we view **mono-chromatic** light with the spectroscope, that is, a light composed of but one color, we see only a vertical image of the slit in B, and its position will depend upon the refrangibility of that color. If, on the other hand, we illuminate the slit with a light containing several rays of different degrees of refrangibility, we shall see one image for each ray present; and they will be separated from one another by their differences of refrangibility. The same color always appears in the same position with reference to the others. If we look at a solid body, heated till it emits a pure white light, there will be so many images spread out on the field of vision, that the one overlaps the other until there are no dark spaces between them, and thus gives a **continuous spectrum**, as it is called. If we place a light before the slit which emits very little light of its own, such as that given by a Bunsen burner, and then put into this flame a little sodium, which gives a pure yellow light, we shall see but one image of the slit, and that in the position occupied by the yellow, in the continuous spectrum, or the D line in the solar spectrum. If we use lithium instead of sodium, we get one image in the position occupied by the red; if thallium be used, the single image is seen in the position of the green. If we mix the three, we shall see the three images as bright lines, each in its own position; one in the yellow, one in the red, and one in the green. In the better instruments, the yellow sodium line appears as two parallel lines. If we illuminate the slit with the vapor of a metal which emits rays of several different degrees of refrangibility, we shall see several images or bright lines in different portions of the field. Any given element always emits the same rays under like conditions; hence, by use of the spectroscope, we may determine what element is introduced into the flame of the lamp by the lines we see in the telescope.

39. **Solar Spectrum—Dark Bands.**—If we illuminate the

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* For an explanation of the principle of lenses the student is referred to works on Physics.

slit with the light of the sun, we see almost a continuous spectrum marked by a number of dark lines, known as "Fraunhofer's lines," the cause of which we shall refer to again.

The most prominent of these dark lines have been designated by the letters of the alphabet, as will be seen on reference to the solar spectrum in the **frontispiece**. These lines serve as landmarks upon the spectrum, by which we can fix the position of other lines, or by which we can designate the position of any line. The D line, for example, is the most brilliant, and can always be seen in the solar spectrum. This serves as a starting point in mapping the spectrum, or as a guide in focusing or adjusting the telescope.

40. **Spectrum Analysis.**—The spectroscope is an important aid to chemical analysis when used with certain precautions, and with certain well-known facts, in its use, kept in mind.

We shall state briefly the principles upon which spectrum analysis is founded:—

1st. All bodies, when intensely heated, become luminous, and, other things being equal, the higher the temperature the more intense the light.

2d. Solid and liquid bodies, if opaque, emit when heated, first a red light, but as the temperature rises the other colors make their appearance, and finally predominate. If the temperature reaches what is called a blue heat, the blue and violet begin to predominate.

3d. The elementary substances give their characteristic and peculiar light only in the state of gas or vapor. Hence, when we examine the light from any given source, we may conclude from a continuous spectrum that the heated substance is a liquid or solid, while a broken spectrum is due to a luminous gas or vapor; and, from the position of the bright lines, determines the nature of the substance giving the light. There are, however, a few exceptions to this principle. Under certain conditions even a gas may give off a light that will give a continuous spectrum.

4th. At the temperature at which gases or vapors become luminous, compound bodies, as a rule, break up into their constituents, *i. e.*, the elemental atoms seem to be dissociated. For this reason, little is known of the spectra of compound bodies.

5th. At a high temperature the metallic atoms are much more luminous than the non-metallic ones with which they are associated. Hence, when we examine the vapor of a metallic salt rendered luminous, the light emitted is so exclusively that of the dissociated metallic atoms, that, whatever salt of that metal be used, we obtain essentially the spectrum of the metal itself.

6th. If, when the slit of the spectroscope is directed towards a source of white light giving a continuous spectrum, another flame giving a mono-chromatic light from a luminous vapor be interposed between the white light and the slit, a dark image of the slit will appear in the position where the vapor of itself would have given a bright line. That is, when the light from a liquid or solid luminous body is made to pass through a luminous vapor, those rays of light are absorbed which the vapor, of itself, emits. Hence, when we analyze the light of a distant source of light, and observe a continuous spectrum marked by dark lines, we may conclude that it is produced by a solid or liquid luminous body shining through a luminous vapor. This explains the dark lines seen in the solar spectrum.

7th. Many substances in solution absorb certain rays from a beam of white light passed through them, and the portions of the beam absorbed are peculiar to each substance. We thus have a means of detecting the presence of a few substances which cannot be rendered luminous, by passing a white light through the solution suspected to contain them.

41. **Absorption Spectra.**—When we wish to observe the spectrum of a liquid we place it in a glass tube, or preferably a vessel having parallel sides, and place this before the slit of the spectroscope. We now throw a strong white light through the solution and into the slit of the instrument. Solutions of **erbium** and **didymium** examined in this way absorb certain portions of the spectrum given by the source of light. The particular portions of light absorbed are peculiar to each, and in these cases the dark bands across the bright spectrum occupy the same positions in which the vapors of these elements give light bands. Absorption bands differ from the dark lines of the solar spectrum in being broader and not so sharply marked. They are often only a slightly darkened portion of a bright spectrum. Even passing the light through a crystal, in some cases, gives the same result as passing it through its solution.

The use of the spectroscope in medicine and toxicology is chiefly confined to the observation of absorption spectra of various solutions. Some idea of the appearance of such spectra may be obtained by reference to the frontispiece, remembering that the spaces which appear white there are, in practice, occupied by the colors of the solar spectrum in their appropriate places.*

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* See Rosenberg or McMunn on the use of the spectroscope in medicine.

42. **Chemical Effects of Light.**—If a mixture of pure hydrogen and chlorine gases be prepared in the dark, and kept there, no combination takes place; if the mixture be brought out into a light room, a gradual combination takes place and hydrochloric acid is the result; if the mixture be placed in the direct rays of the sun, instead of diffused light, the combination takes place with an explosion. The light in this case causes chemical action. The electric light and other intense lights produce the same action. If a piece of white paper wet with a solution of nitrate of silver be kept in a dark room no change takes place in it; but if the paper is exposed to a strong light for a few minutes, it begins to grow dark and finally becomes black.

Many chemicals kept in the light are in time sensibly changed. Silver and gold solutions in presence of organic matter deposit a part of their metal in the metallic form. Nitric acid becomes slowly yellow, from the decomposition produced by the light. A solution of the syrup of the iodide of iron on exposure to air becomes brown from decomposition; but on exposure to the direct sunlight it again recombines and returns to the normal green color. The art of photography is based upon the changes produced in silver salts, gelatine, etc., by light. Sometimes the change seems to be a true chemical reaction, and in other cases it is a molecular change only.

It has been found that it is not always the luminous part of the ray of light that effects these changes, but that they are chiefly produced by certain invisible rays found most abundantly beyond the violet part of the spectrum. From this it has been concluded that the light of the sun, as well as the light from some other sources, contains certain rays having this power to produce chemical changes. This action is known as **actinism**, and the rays producing it are called **actinic rays.**

ELECTRICITY.

43. **Electricity Produced by Chemical Action.**—If in a vessel of water, containing a little sulphuric or hydrochloric acid (1 to 20), a strip of zinc and one of copper or platinum be immersed and prevented from coming in contact, no action is seen to take place, provided the acid and zinc be pure. If, however, we connect the two strips of metal by a wire, or the upper ends of the strips are brought in contact above the liquid, chemical action immediately takes place, and the following phenomena are observed: 1st. Small bubbles of gas are seen to collect on the surface of the platinum strip, while the zinc slowly dissolves, and at the same time the acid begins to disappear. In

the case of hydrochloric acid, the chlorine combines with the zinc, and the hydrogen escapes from the opposite plate. 2d. We shall find a peculiar property manifested by the wire, which is known as electricity. If a magnetic needle be placed near the wire it is turned from its course. If the wire be broken and the tongue be placed between the two ends, a tingling, metallic taste is observed. 3d. If the plates are large, and the ends of the wires are placed near together in a solution of copper sulphate, the metallic copper begins to deposit on one of the wires. In a word, a force is developed in the wires which we call **electricity.** If the wires are separated from each other by air the chemical action ceases, the gas ceases to escape, and the zinc to dissolve. The same phenomena are observed when we substitute for the above metals zinc and lead, zinc and gas retort carbon, etc. It is only necessary that the plates be unequally acted upon by the fluid in which they are dipped; and the greater this difference, within certain limits, the stronger is the force developed in the wire. In order that these phenomena shall take place, the following conditions are necessary:—

The plates and connecting wires must be conductors of electricity.

The liquid must contain some substance with which one of the plates can form a chemical action.

44. **Electrical Induction.**—By this term is meant the property an electrified body possesses, of inducing or causing an opposite condition in certain bodies near it. This is a property of electricity that we must understand before we can comprehend the action of the galvanic cell.

If a magnet is brought near any piece of iron, as a nail, the latter becomes a magnet, with the poles in the same position as those of the original magnet.

So, a piece of brass held near the prime conductor of a friction electrical machine, becomes charged with the opposite kind of electricity to that of the conductor. A negatively electrified cloud passing over the earth induces the opposite condition in objects directly under it on the earth.

This condition of a body is generally known as polarity; *i. e.*, the body manifests a peculiar force at certain points known as its poles. A magnet usually has two poles.

The chief characteristics of this polar force are as follows:—

1st. The energy is chiefly exhibited at opposite points of the body, called its poles.

2d. Like poles repel and unlike poles attract each other; and while like poles enhance each other's effect, unlike poles neutralize each other.

3d. A polarized body induces a like condition in all bodies near it susceptible of this condition, a pole of a given kind inducing a pole of the opposite kind in that part of the body nearest to itself.

4th. Induction is attended with no loss of energy in the inducing body, but on the other hand its polarity is rather increased by the reaction of the induction.

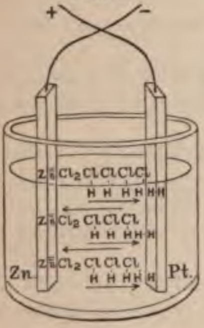
5th. Magnetic, electrical, and chemical polarity are somewhat similar in their general features, but differ considerably in their modes of manifestation.

6th. Bodies differ not only in their susceptibility to polarity, but also in their power of retaining it. Most atoms have an inherent chemical polarity which is more or less permanent and characteristic, although it may be increased, diminished, or entirely changed in quality, by induction.

Chemism or chemical affinity is the manifestation of this polar energy acting between two differently polarized atoms. Every molecule may be regarded as naturally polarized, one part being positive while the other is negative.

45. **Theory of the Galvanic Cell.**—In order to bring the working of the cell clearly before the mind, let us assume the plates in Fig. 9 to be platinum and zinc, and the exciting fluid to be a solution of hydrochloric acid. The space between the plates may be considered to be filled with molecules of the acid, each molecule composed of one atom of negative chlorine united to one atom of positive hydrogen. The negative chlorine atoms have an inherent tendency to combine with the positive zinc atoms, which are capable of taking a higher positive condition than hydrogen, and by this tendency the zinc atoms become strongly charged with positive electricity or positive polarity, which they retain as long as they are in combination with the chlorine. By this means the rest of the zinc plate becomes negatively polarized or charged with negative electricity.

Fig. 9.



Now, it is clear that if there is no chance for the plate to communicate with the earth or some neutralizing body, it will soon reach the maximum degree of polarity that can be induced by the chlorine. The action must then cease until the equilibrium of the polarity or electricity in the plate has been established; or, in our primitive cell, until the wires connecting the plates are brought in contact, when a neutralization takes place, and the action continues.

Thus far, we have considered only the negative plate. When the chlorine atoms leave the hydrogen for the more positive zinc atoms, the hydrogen seizes upon the neighboring chlorine atoms in the adjoining molecules, and thus a stream of hydrogen atoms sets toward the platinum plate and are finally liberated at its surface. At this point two atoms combine to form a molecule of hydrogen; but in order to do so, one of the atoms must discharge a part of its positive charge of electricity upon the platinum plate, thus charging it with positive electricity.

By the above process, the zinc plate is rendered negative, or its electric tension is lowered below the normal by the withdrawal of positive electricity, while the platinum plate is actually charged above the normal, and is in a positive electrical condition. It is this difference of electrical state which causes a neutralization through the wires as long as the exciting fluid and the zinc plate last, or as long as the chemical action continues. If the wires are disconnected, the platinum plate soon becomes charged to a tension equal to that of the hydrogen atoms, and the zinc plate lowers its tension until the chlorine atoms will no longer leave the hydrogen to combine with it. The action then ceases until the connection is again made. The discharge is then continuous from the platinum to the zinc plate, and it can also be shown that a current passes through the liquid. Such is the best conception we possess of the simple galvanic cell which we have considered.

46. **The Current or Circuit.**—The circuit is said to be **closed** when the wires are connected, and there is a constant flow or transfer of force through the wires and through the liquid. It is said to be **open** or **broken** when the wires are separated so that the transfer of force ceases.

If the wires from a galvanic cell or a collection of cells be connected with the earth instead of with each other, the current flows as if the latter were really done; and it makes no difference how far apart the wires connect with the earth. This is called **grounding** the battery. No current actually flows from the one point to the other in the ground, but by bringing the plates in contact with the earth their electrical equilibrium is restored by it. This principle is made use of in telegraphy to avoid the necessity of a return wire. One of the wires of a battery situated at one of the stations is **grounded**; while the other passes to and through the other station and is then grounded at that point; so that the current must pass between and through both stations to complete the circuit.

47. **Electrical Tension or Electro-motor Force.**—When we speak of the normal electrical condition, we have reference to the electrical state of the earth or bodies in contact with it. The earth is the great storehouse of electricity, as the ocean is of water. If water is taken up from the ocean, and deposited upon the mountain side, it will run back to the sea in a stream, and can be made to do work, while reaching its former level, by turning a water wheel. The water, in other words, has acquired power to do work by its change of position, which in mechanics, is called its **potential.** In the galvanic cell, the equilibrium of the electricity is disturbed, and it acquires power to do work in returning to its former state of electrical equilibrium. This property in electricity is called **tension** or **electro-motive force** (E. M. F.). The strength of the E. M. F. will, of course, depend upon the difference in the electrical condition of the two plates. It is our purpose to consider here only such practical points as we deem essential to the physician's knowledge in the use of galvanic batteries.

48. **Forms of Cells.**—We have thus far discussed only one form of cell. The term **battery**, strictly speaking, is applied to a collection of cells; but it is frequently applied to a certain form of cell. Various kinds of cells are in common use.

One difficulty in the working of the simple cell we have already described, is in the fact that the hydrogen accumulates on the platinum plate and prevents contact with the liquid, and thus obstructs the current. In order to obviate this, various means have been used to prevent this gas from reaching the platinum.

In **Grove's** cell, the platinum plate is suspended in a porous earthenware cup filled with strong nitric acid, and placed in the centre of the larger cup, containing the dilute sulphuric acid (1 to 12). The nitric acid oxidizes the hydrogen, converting it into water before it reaches the platinum.

Bunsen's cell is constructed in the same way as the above, except that the platinum is replaced with the cheaper gas retort carbon.

In the working of the above cells the nitrous fumes evolved are very objectionable, and to avoid this, a solution of chromic acid in sulphuric acid, made by adding to 18 parts of water 4 parts of potassium bichromate and 4 of sulphuric acid, may be used. The chromic acid serves to destroy the hydrogen in the same way as the nitric acid, and no porous cup is needed. The elements used are zinc and carbon. This cell gives a strong current for a short time, and is one of the best in use, for medical purposes. The zinc plates are always removed from the liquid when the battery is not in use. In some medical batteries a solution of acid sulphate of mercury in water is used as the exciting fluid, instead of the above. In this case the plates are small, and made of zinc and carbon. The zinc combines with the sulphuric acid, and mercury instead of hydrogen is set free. Another form of battery is one in which the exciting fluid is dilute sulphuric acid and the elements zinc and silver, the latter being inclosed in a layer of chloride of silver, which is intended to prevent the hydrogen from accumulating on the silver plate, by combining it with chlorine. $2AgCl + H_2 = Ag_2 + 2HCl$. These cells are usually made in the form of long, narrow cylinders, so as to occupy a small space, and are very constant and effective.

In the **Leclanché** cells, as usually constructed, the elements are a plate of carbon and a rod of zinc. The carbon plate is surrounded by a layer of peroxide of manganese, to serve as a depolarizing agent, or in a slightly different form, ferric oxide is used for the same purpose. The exciting fluid is a strong solution of ammonium chloride in water. This battery is one in very common use where an open circuit is to be used, is very constant, requires attention only at long intervals, and is inexpensive. Various modifications of this cell have been proposed, but in efficiency they are not superior to it.

The **Callaud** battery is constructed as follows: The elements are zinc and copper. The former is suspended in the upper portion of a solution of copper sulphate contained in a glass jar. The copper plate lies at the bottom of the jar, and the wire attached to it is covered with gutta percha for the purpose of insulating it. From time to time copper sulphate crystals are dropped into the jar, to keep up the supply.

This battery is useful where a closed circuit is to be used, and the battery is to be in constant use. It is very constant when kept in good order, but has a low electro-motive force, and is seldom used in the construction of medical batteries. A new form of cell, which is said to give good results, is made as follows: An iron tray or pot is nearly filled with a strong solution of ammonium chloride, and a plate of zinc is supported by non-conducting posts in this solution. A thin layer of oxide of copper is sprinkled over the bottom of the tray, to act as a depolarizer. The wires are attached to the zinc and iron tray. When the oxide of copper becomes reduced, and loses its effect, it may be again oxidized by roasting it in the air.

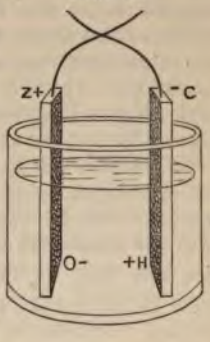
49. **Care of Batteries.**—In order that a battery may perform its work it will need some care in its management. All metallic connections, as well as the wires through which the current is to pass, must be of good conducting material. Copper or silver wire is usually employed for conductors, and where two wires are meant to connect, their surfaces must be bright and free from oxides, which are poor conductors. As far as possible, a uniform strength of exciting fluid should be maintained. In most batteries this will require entire renewal, from time to time, in order to supply new material for chemical action, and to remove the products of former action.

50. **Local Currents.**—Owing to the imperfections in the zinc used in the construction of batteries, it is unequally acted upon by the liquid. The points where the zinc is harder, or contains iron, lead, or arsenic, act as negative plates to the rest of the zinc, and thus currents are set up between them which eat away the zinc, and cause a serious loss of material, as well as of force. "Amalgamation" of the zinc obviates this action by forming over the surface of the plate a homogeneous layer of zinc amalgam.

To amalgamate the zincs, first wash them in dilute sulphuric acid (1 to 6), and then pour mercury upon them while still wet and rub in the drops until the whole surface is uniformly bright and smooth. It is well to keep a little mercury in the bottom of each cell, which keeps the plates amalgamated. A hissing sound, or the evolution of hydrogen from the surface of the zinc, is a sure sign that the zincs need re-amalgamating.

51. **Polarity of the Elements of Batteries.**—A serious hindrance to the working of batteries is what is called the **polarization of the plates.** We have already referred, when speaking of the construction of cells, to the accumulation of hydrogen on the carbon or platinum plate. When the current is of considerable strength, oxygen accumulates on the zinc plate and hydrogen on the opposite one. We then have a layer or plate of hydrogen against the carbon, and a layer or plate of oxygen against the zinc. The former of these is positive and the latter negative; and they are joined together by the same

FIG. 10.



wires as the primary plates, as will be seen by a glance at Fig. 10. Not only is the liquid kept from perfect contact with the plates, but, owing to the action of the liquid upon these new gaseous plates, a current is developed in the opposite direction to that of the primary current, which may become so strong that it almost entirely overcomes the original current and destroys the efficiency of the battery.

Some method must, therefore, be adopted to prevent the hydrogen from accumulating upon the carbon or platinum plate. Nitric or chromic acids, oxides of manganese, copper or iron, silver chloride, and copper sulphate are all used for this purpose, as referred to in article 48.

52. **Secondary or Storage Batteries.**—The polarity of the plates of a battery cell is utilized in the **secondary or storage batteries.** The cell contains two or more plates of large size, constructed of sheet lead, or finely divided lead pressed into plates, and is filled with dilute sulphuric acid. The plates are polarized by passing a current through the battery. The hydrogen accumulates in or upon one plate, and the oxygen in the other. On now disconnecting the charging battery, it is found that a current may be obtained from the polarized cell for some time, but in the direction opposed to that of the charging current. When the plates of this battery are once charged, they will remain charged for some weeks; and the current may be obtained at any time, whenever the wires from the opposite plates are connected.

53. **Resistance of Conductors.**—Conductors are bodies which allow a ready transmission of the electrical impulse through them, and are contrasted with another class of bodies called non-conductors or insulators. These terms are only relative, however.

Some bodies conduct electricity with great ease, while others offer more resistance to the passage of the current, or entirely refuse to allow an appreciable amount to pass. Even the best conductors offer some resistance to the passage of the current. The metals are the best conductors, and of these silver is the best conductor known. Copper is second to silver only, and when both metals are pure the difference is but slight.

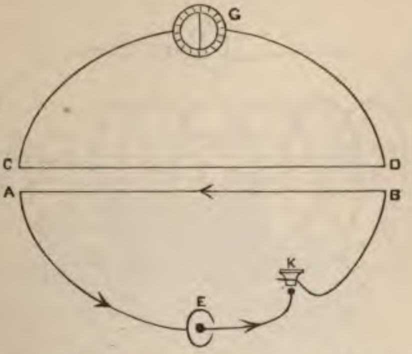
If we compare wires of the same material, but of different sizes and lengths, we find that **the resistance of wires increases with the length, and diminishes with the area of the cross section.** When a cell is in action, the current not only meets with resistance in the wires, but also in the liquid of the cell through which it has to pass. This last resistance is usually much greater than that of the wires, and is an important element in determining the strength and intensity of the current.

54. **Ohm's Law.**—This law states that the intensity of a current developed by a battery will be equal to the electro-motive force divided by the resistance. By electro-motive force we mean the force with which the electric current is set in motion. This law may be stated algebraically as follows: $C = \frac{E}{R + r}$, where R represents the internal resistance of the liquid, r the external resistance or that offered by the wire, and E the electro-motive force, which is always the same when the same metals and exciting liquid are used. In any given form of battery, variations in the strength of the current must be due to the variations in resistance, either in the external or internal part of the circuit, or to a change in the strength of the exciting liquid, polarity of the plates, or secondary currents. We have already spoken of the resistance offered to the current by the polarity current, flowing in the opposite direction, and which may sometimes become almost as great as the electro-motive force can overcome. It is clear that in order to increase the value of C, in our formula, we can increase E or diminish R and r. To increase the electro-motive force we select such metals and liquids as shall give us a relatively high intensity of current. We may increase the intensity of the E. M. F. by joining several cells, so that the force of the one may be reinforced by the next, and so on. This is done by connecting the zinc of the first to the carbon of the second, the zinc of the second to the carbon of the third, etc. Each cell added to the series adds to the current its E. M. F. diminished by its internal resistance; the external resistance being too small to be regarded. The formula applied to the series would be, when n equals the number of cells: $C = \frac{n E}{n R + r}$. Now when the external resistance in the wire, r, is very small in comparsion with R, as when flowing through an ordinary copper wire, it may be disregarded; and the equation then becomes $C = \frac{E}{R}$. That is, the effect of a battery of several elements in this case is no greater than that of a single cell. If, however, the external resistance, r, is great, as when the electrodes are applied to a human body, which has a resistance many times

greater than the usual value of R, the value of C increases or diminishes very nearly in the same ratio as the number of cells. For medical purposes, therefore, we usually combine the cells as above described. Elements or cells so arranged, are said to be arranged **in series**, or arranged **for intensity**.

We may also increase the value of C, in the formula above, by increasing the size of the plates, provided the external resistance is small. By so doing, we do not increase the electro-motive force, but we reduce the resistance in the cell by virtually combining several plates into one and increasing the surface exposed to the liquid, without increasing the distance through which the current has to pass in the liquid. Where a small resistance is to be overcome, therefore, large plates are to be preferred; or, which is the same thing, all the zinc plates of the battery may be connected, and all the carbon plates. When the cells are arranged in this manner, they are said to be arranged in **multiple arc**, or **for quantity**. The **poles** or **electrodes** of a battery of cells are the conducting wires; that attached to the zinc plate is the **cathode** or **negative** electrode, and that attached to the platinum, carbon, or copper plate is the **anode** or **positive** pole.

Fig. 11.

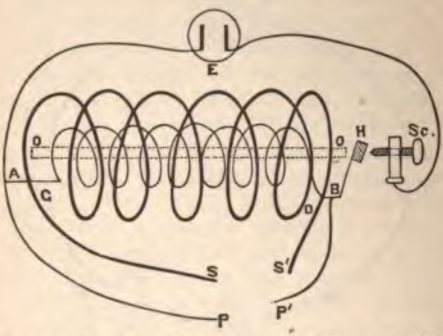


In electro-therapy the term **electrode** is often used to designate the appliance fastened to the end of the wires for application to the patient, while the wires are called **rheophores**.

55. **Induced Currents.**—If a current of electricity, from a battery, be passed through one of two parallel wires A B (Fig. 11), lying near together, no current is observed in C D as long as

the current in A B is constant; but if this be abruptly stopped, an instantaneous current is developed in C D, which we can demonstrate by connecting the wire to a galvanometer (G). When we make the current pass from B to A, the current in the wire C D takes the direction from C to D; but on breaking the primary current, the induced current takes the direction D to C. If, therefore, we rapidly make and break the primary or battery current by means of the key, K, we shall have a rapid to and fro current acting in the secondary wire C D. Now, if these wires be covered with silk or insulated, and are wound together around a spool or bobbin, the conditions of the experiment will remain unchanged, and we shall have the same phenomena in the coiled wires as if they were straight. Such a coil is known as an induction or Rhumkorff coil. The strength of the current in the secondary wire, or the induced current, will depend upon the length of the wire acted upon, the strength of the battery or primary current, and the distance the wires are from each other.

FIG. 12.

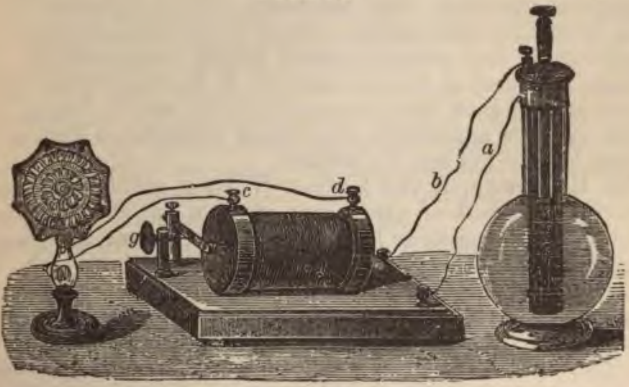


56. **Induction Coil.**—It is customary, in constructing an induction coil, to make the primary coil of large, thick wire, so as to allow the battery current to pass with as little resistance as possible, and to make the secondary coil of a much longer and thinner wire. The former is made into a smaller coil, which slips into the latter. Into the inner coil is pushed a bundle of soft iron wires which act as magnets when the battery current is sent through the coil. A small armature or piece of soft iron fastened to a spring, vibrates before the end of the bundle of wires. When no current is passing the spring rests against the point of the screw Sc, Fig. 12. When a current is sent through

from B to A, in the inner coil of wire, an induced current is produced in the outer coil, from S to S′, or from C to D. At the same instant the current B A magnetizes the core of wires O, and the hammer H is drawn towards them and away from the point of the screw. This breaks the current at that point, the core de-magnetizes, and the spring brings the hammer back to the screw, when the process is repeated, as long as the current from the battery E lasts. The induced current in S S′ is, therefore, a **to and fro** current, or a **make** induced in one direction and a **break** induced in the opposite direction.

This current is known as the **Secondary, Induced, Interrupted** or **Faradic Current.** The two wires *a* and *b*, Fig. 13, are connected with the primary coil, by the binding posts, and carry the battery current. The wires *c d* are the terminal wires of the outer secondary coil, and carry the induced current. The interrupter is shown at *g*.

FIG. 13.



57. **Extra Current.**—It is very evident that each turn of the wire in the primary coil lies very close to and parallel with the adjoining turns of the same wire, and that these consecutive turns may be considered as constituting a series of parallel wires. In fact, every variation of the current in the wire A B Fig. 12 generates electro-motive force in the contiguous turns. An induced current is thus produced in the wire A B, which obeys the same laws as that induced in the independent wire, C D, and in the direction opposed to the

battery current, when the latter is made or increased, and in the same direction when it is broken. This current is known as the **primary induced** or **extra current.** During the making or increasing of the battery current this **extra current,** acting against the battery current, retards or resists it, and hence is not felt at the poles P and P′. At the **break** this current goes in the opposite direction, and, as there is nothing to resist it, may be felt with its full force at P and P′. This current is therefore **interrupted,** and is felt only at the break of the battery current, and always in one direction. The primary induced, or extra current is feebler than the secondary, because the length of wire acted upon is shorter. As we have just seen, the **make** extra current retards the battery current, so that it takes time for this current to attain its maximum force; and the **make** induced current is weakened in proportion to the longer time required. The secondary current developed at the time of making the battery current is therefore weak, and its physiological and chemical effects almost inappreciable. The **break,** secondary as well as primary, is developed with its full electro-motive force instantaneously, so to speak; hence, they alone have an appreciable effect when a resistance such as the human body is put into the circuit.

58. **Influence of the Core.**—When a galvanic current is sent through a coil of wire wound about a bar of soft iron, the bar becomes a magnet as long as the current passes, and loses its magnetism as soon as the current in the wire is broken. The effect of the bundle of soft iron wires is the same as that of a single bar. Moreover, when a magnet is suddenly made or destroyed, it causes a current to flow through the wire wound about it. The effect of the magnetic core is, then, only to retard the battery current when it is first passed through the coil, and to still further weaken the induced currents developed by it. Its sudden de-magnetization reinforces the break currents and makes them stronger. The currents are further modified by means of a draw tube made to inclose more or less of the primary coil. When this is completely withdrawn the current is strongest; and when the inner coil is completely enclosed by it the currents are considerably weakened. Occasionally the secondary coil is made to include any desired length of the primary, and thus the current may be varied at will. Figure 13 shows an induction coil with Grenet cell ready for use.

59. **Magneto-Electricity.**—Besides chemical action, other methods of producing electricity may be employed. We have

already referred, in the last section, to the effect of suddenly making and destroying a magnet within a coil of wire. The same effect is produced when the magnet is made to approach or recede from the coil of wire, or when the magnet is increased and decreased in strength. The simplest magneto-electric apparatus is composed of a strong horseshoe magnet, before the poles of which two short soft iron bars, called armatures, mounted on a shaft and wound with coils of wire, are made to revolve by a crank. As the soft iron approaches the poles it becomes a magnet and induces a current in the wire. As it recedes from the same pole the current in the wire is reversed, but by a pole changer it is made to pass from the instrument in the same direction.* By multiplying the number of magnets and armatures, by using the current developed to strengthen the magnets, and by revolving the armatures at a high rate of speed, a very large amount of electricity may be produced. The electric light is now very generally supplied with electricity by machines made upon this principle and driven by steam or water power. Such machines are called dynamo-electric machines. Small machines are sometimes employed for medical purposes, which have been improved by Mr. Gaiffe so as to give much better results than the older forms of this apparatus. In these machines the electricity is the result of the conversion of mechanical force into electrical force, the two being mutually convertible.

60. **Thermo-Electric Currents.**—If two bars of dissimilar metals be soldered together at one end, thus: $\underset{c}{\overset{a}{\diagdown}\overset{b}{\diagup}}$, and the junction *c* be heated while the ends *a* and *b* are cool, electricity will pass; the direction of the current will depend upon the metals composing the couple. If the metals be bismuth and silver, the current will be from the former to the latter; if German silver and iron, the current will be from the former to the latter. By arranging a large number of such pairs in a pile, thus:—$-\diagdown\diagup\diagdown\diagup\diagdown\diagup\diagdown\diagup+$, so that alternate junctions can be heated, a current of considerable strength may be produced. Such a pile may be constructed by arranging the bars in the form of rays around a hollow centre, in which a lamp or fire can be kept burning, and thus furnish a constant current of electricity. Thermo-electrical currents have not been used in therapy, but it

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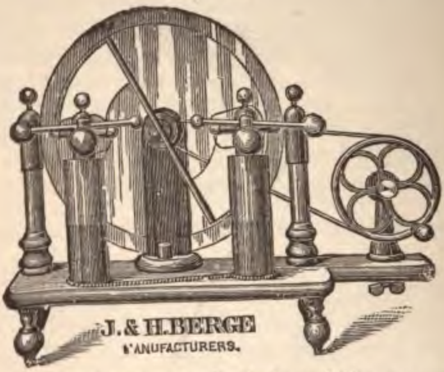
* In another form of the machine the coil of wire is on the magnet, and the current is produced by the variations in the magnet itself, caused by the armatures.

seems probable that a pile properly arranged may be made to furnish a current in every way equal to a galvanic current.

61. **Frictional or Static Electricity.**—When a dry glass rod or tube is rubbed briskly with a piece of dry flannel, silk, chamois leather, or cat's fur, for a few seconds, it acquires the property of attracting light objects, as feathers, dust, or bits of paper to its surface, and is said to be electrified. Sealing wax, resins, gutta percha, etc., exhibit the same property to a greater or less extent.

After the feather, or other light body, has adhered to the glass or sealing wax for a few seconds, it falls off; if the glass is now excited anew, and presented to the feather, it repels it. Or, if the feather be first made to repel from the glass, it will be found to be attracted by the sealing-wax with much greater force than before. From this it will be seen that there is a parallelism between this kind of electrical excitement, and magnetism; the excited glass and sealing wax bearing the same relation to each other as do the north and south poles of the magnet.

FIG. 14.



Machines for developing statical electricity are usually made of a glass disc mounted on a horizontal axle, and turned by a crank. Friction is caused by two rubbers, covered with amalgam, one on each side of the disc, regulated by a screw, and connected by a chain with the earth. The electrical condition of the glass plate is disturbed by the friction, and the positive electricity is removed from the plate by a row of small points connected with the prime conductor, and brought near the revolving plate. In the Holtz machine, which is the one now generally used, the rubbers are dispensed with, and the electricity developed by the induction of a body previously electrified. (See Fig. 14.) For a detailed description of this machine the student is referred to works on physics. Static electricity is capable of giving all the effects of galvanic or Faradic currents, when properly managed. It may also be used instead of the electric bath.

62. **Physiological Effects of Electricity.**—Electric currents exert a marked physiological effect upon nerves, muscles, and the circulation of the human body, when they are made a part of the circuit. A constant or galvanic current has a refreshing effect upon a nerve through which it is passed, as well as upon a muscle, the latter being able to perform more work under its influence than without it. Under its influence the circulation of the blood may be increased in any part of the body, which lasts for some time after the electrodes have been withdrawn. The physiological effect of the negative pole seems to be greater than that of the positive. The interrupted, induced, or Faradaic current stimulates nerves most when passed in the direction of the natural nerve current; this seems to be its principal physiological effect.

For more extended remarks upon this part of the subject the student is referred to one or more of the manuals* upon the subject.

63. **Chemical Effects of Electric Currents.**—When a strong galvanic current is passed through a vessel of water containing a little sulphuric acid, the liquid is decomposed, hydrogen gas is given off at the negative pole, and oxygen at the positive. This process is called **electrolysis.** If we perform the same experiment with a solution of a salt of one of the metals, the metal appears at the negative, and the negative element or radical appears at the positive pole. If the same current be passed through an animal tissue, the following changes take place: The water in the tissue is decomposed, the hydrogen appears at the negative pole, along with the metals of the salts, while the oxygen, with the non-metals or acid radicals, appear at the positive. The nascent oxygen surrounding the positive pole attacks the neighboring tissues and converts them into a hard eschar, while the alkalies at the other pole exercise their caustic properties and form a soft, frothy mass, containing hydrogen gas in small bubbles.

When a sufficiently strong and somewhat prolonged current is applied to the skin a similar effect to the above is produced at the point of contact of the electrodes. At the positive electrode the skin first becomes red, with a burning sensation, then an eschar is produced with an acid reaction of the tissues. The eschar resembles that produced by a strong acid. At the negative electrode there is a vesicle formed with an alkaline liquid; if the action be prolonged more extensive ulceration takes place.* From this it will be seen that strong currents should be used with care. The effects of the magneto-electric current are very similar to if not identical with those of the galvanic current. Recently the use of frictional electricity in the treatment of disease has been revived and advised as a substitute for Faradic electricity. The effects of this form of electricity have not been so well studied as those of the Faradic and galvanic, but clinical results seem to indicate that its effects are even more pronounced than those of the latter.

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* De Watteville, Bartholow, or Beard and Rockwell.

64. **Chemical Effects of the Galvanic Current Outside of the Body.**—We have seen that when a current of electricity from a battery or a magneto-electric machine is passed through a solution of the salts of the metals, they are decomposed, the metal appearing at the negative, and the negative radical at the positive pole.

The salts of some of the metals—the earths and alkalies—require a very strong current, while some of the other metals do not require more than the current of one or two ordinary Leclanché cups. If the current used be not too strong, the metal is deposited upon the negative electrode in a compact, tenacious form, adheres very firmly to the surface, and is capable of taking a polish. For this reason, the process of depositing metals upon the surface of other metals has come into extensive use in the arts of electro-plating and electrotyping. The principal metals used to deposit upon others, in this way, are gold, silver, copper and nickel. The objects to be attained are to protect easily oxidizable metals from rust, to preserve a brilliant surface, and to coat cheaper metals with the more valuable ones. This process is known as **electro-plating.** Only the outlines of these operations can be described here.

65. **Electro-Metallurgy or Electro-plating.**—Silver and copper are more easily deposited than most other metals. The strength of current needed to deposit these metals is rather feeble, unless the surface to be coated is large. The quantity of electricity should be varied according to the surface to be coated; larger surfaces requiring a stronger current than smaller ones.

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* De Watteville, pp. 33 and 34.

The strength of the solution of the metal to be deposited, will be governed somewhat by the material composing the article to be plated and the strength of the current, and will have to be determined experimentally by the beginner. When the proper strength of solution, current, etc., have been found, these should be kept as nearly constant as possible. The plating solution may be kept of constant strength by suspending from the positive pole of the battery a plate of the same metal contained in the solution, the size of which should be nearly equal to the size of the article to be plated.

66. **Preparation of the Surface to be Coated.**—The article to be plated must be thoroughly cleansed from rust and grease, and then, without handling with the fingers, suspended in the solution, taking care that no bubbles of air are left clinging to its surface. It should be suspended from the negative pole (that coming from the zinc of the battery). To insure thorough cleansing, the article is to be polished as free as possible from rust, dipped in dilute sulphuric or hydrochloric acid, thoroughly washed, and dipped into a solution of caustic soda, potash, or ammonia, to remove grease, again washed, and put into the plating fluid without drying or handling. The current is immediately connected and allowed to pass for a few minutes, or until a complete coating has been deposited. The article is then removed and burnished with a soft brush and a little whiting or precipitated chalk, thoroughly cleansed, and returned to the solution, and the deposition allowed to proceed with one or two repetitions of the burnishing process. In the case of silver, the labor of burnishing may be lessened somewhat by the addition of 4 or 5 drops of bisulphide of carbon to each pint of the plating solution.

67. **Preparation of the Solutions.**—The composition of the solutions varies somewhat, according to the kind of surface to be plated.

The Copper Solution used to deposit upon copper, brass, white metal, and all metals which are more electro-negative than itself (See Art. 86), is a solution of copper sulphate in water. If a more electro-positive metal, like iron, lead, or zinc is to be plated, the double cyanide of copper and potassium is to be used. If a strip of clean iron be plunged into a solution of copper sulphate, a brown-red non-adherent deposit immediately takes place, without the aid of the current. It is therefore necessary to employ a solvent for the copper which shall have a different relative affinity for the copper and iron from that which the acids have. Such a solution is the one mentioned. The solution may be prepared as follows: 1. Take a solution of copper sulphate, and add to it a solution of potassium cyanide or potassium ferrocyanide until a precipitate ceases to form. Allow the precipitate to settle, and pour off the clear liquid. Add some water, allow it to settle, and again pour off. Repeat this process three or four times. Then pour upon the precipitate a solution of potass. cyanide (2 lbs. to 1 gal. water), until the precipitate dissolves; it is then ready for use. 2. Or, we may suspend the article to be plated (from a wire of zinc?) in a solution made acid with tartaric acid, and then made very strongly alkaline with an excess of sodium hydrate. Oubry first covers the iron with a layer of some substance which is impervious to acids. He then renders this layer conducting by rubbing it with graphite or plumbago. Results, good.

Silver Solution.—The silver solution generally used is made as follows:—

℞. Potass. cyanide	4 oz.
Silver cyanide	1½ oz.
Water	1 gal.

The Potass. cyanide is dissolved in the water, filtered, if not clear, and the silver cyanide stirred into the solution until it all dissolves. This solution contains 1 oz. of silver in the gallon. Another method: Dissolve a fifty-cent piece in dilute nitric acid. Add a solution of salt until it ceases to produce a precipitate. This precipitates the silver but leaves the copper of the coin in the solution. Now filter, pour off the liquid as described above in the preparation of the copper solution. The moist residue is now to be dissolved in a pint of water in which 1½ oz. of potass. cyanide has been dissolved, when the solution is ready for use. All these solutions should be used at about the temperature of 70° to 80° F.

Gold Solution.—This solution may be made as follows:—

℞. Potass. cyanide	2 oz.
Gold cyanide	55 gr.
Water	1 pint.

The potass. cyanide should be quite pure. The solution should be used hot—from 130° to 150° F. The solution for plating with nickel, is made by dissolving the sulphate of nickel and ammonium in water, in quantity sufficient to give it a grass-green color. It requires a stronger current to deposit nickel and gold than the other metals mentioned. For further instructions in the preparation and use of electro-plating solutions the student is referred to a special work upon this subject.

SOLUTION, DIFFUSION AND DIALYSIS.

68. **Solution.**—The power of water to dissolve substances is one of the most familiar of its properties. All liquids possess the same power to a greater or less extent; but liquids vary greatly in their solvent power, which is usually limited to a certain class of bodies. Thus, mercury will dissolve a number of the metals; alcohol is the proper solvent for the resins, ether for the fats and some gums, and water for the ordinary metallic salts. Water is by far the most universal solvent, and for this reason it is commonly used as the medium of chemical changes.

The solvent power of water varies greatly with different solids. While some bodies absorb water from the air and become liquefied, or **deliquesce,** others require several hundred times their weight of water to dissolve them, and some will not dissolve in it at all. As a general rule, the solvent power of water for solids increases with the temperature; but there are a few exceptions to this rule.

By the solubility of a substance, is meant the amount of the substance which will be taken up by the solvent. The solubility of a substance is absolutely definite at a given temperature, and the amount which 100 parts of water will take has been determined with every known substance.

A knowledge of the solubility of ordinary solids is very important to the medical student as well as to the chemist. The law of compatibles is really the law of solubilities. A table will be found in the Appendix giving the approximate solubility of those salts most commonly met with, and which the student should refer to whenever the question of solubility is spoken of. When a liquid has dissolved all of a solid that it can take up, it is said to be **saturated** at that temperature; but saturation of a liquid with one solid does not prevent it from dissolving others, and, in some cases, the solvent power of the liquid is thereby increased. When two or more salts are dissolved in a liquid, an exchange of the metals and acids may take place, according to the laws of Berthelet, modified by the strength of the affinities of the radicals present.

69. **Solution of Gases.**—Most liquids dissolve gases as well as solids. The quantity of a gas which 1 c.c. of any liquid will dissolve, when the pressure of the gas upon it is 76 cm., is called its coefficient of absorption. As in solids, this coefficient must be determined by experiment for each gas; but, as a rule, this decreases as the temperature increases, although each gas obeys

a rule of its own, which can be determined only by experiment. The **volume** of gas absorbed by a liquid at any given temperature is the same, whatever the pressure. The **quantity** of gas dissolved, therefore, increases and decreases with the pressure. When a liquid is exposed to a mixture of gases, it dissolves each in the same proportion as if it alone were present and exerting its own share of the total pressure. Thus when the air, a mixture of oxygen and nitrogen in the proportion of 1 to 4 respectively, is exposed to a mass of water, we find that the gases are absorbed by the water, in the proportion of 1 to 1.87 respectively.

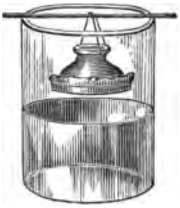
70. **Nature of Solution.**—The term **solution** embraces two entirely different processes. The one a mechanical or **physical**, and the other a **chemical** process. In physical solution the identity of the solid is preserved, as well as that of the water; and by evaporation of the water we may obtain it again unchanged. In some cases there seems to be no manifestation of chemical action between the water and the solid; as when sugar dissolves in it. In other cases, which seem at first sight to be equally simple, there is heat developed, or heat absorbed, which, with other things, leads us to suppose that in these cases there is a true but feeble chemical union of the salt with the solvent. When the solid separates again in crystals it takes a part or all of the water with it as water of crystallization. Simple solution in water may be regarded, in some cases, as a mere overcoming of cohesion of the solid by the adhesion of the liquid to the solid; while in others, it is a feeble combination of the solid with the water, and then the diffusion of this compound through the remaining water. The metallic alloys are in some cases mere mixtures, and in others they seem to be veritable compounds. When a metal dissolves in a dilute acid, there is, at first, a chemical action between the acid and metal, by which a soluble compound is formed. This then dissolves in the water present, as above described. This double process is sometimes termed **chemical solution.**

71. **Diffusion of Liquids.**—When one liquid dissolves in another the process is called **liquid diffusion.** If upon the bottom of a vessel containing pure water, we pour some water colored with a little aniline red, by delivering it through a tube so as to prevent the mingling of the two, and then allow the vessel to remain at rest for some hours, the color will be found to have diffused itself throughout the water.

If instead of a colored wa[illegible] use a strong solution of common salt, having a high specific gravity, we shall find by appropriate tests that the salt has passed throughout the entire liquid. The rate of diffusion in these cases increases, for all substances, with the temperature. This is because the rapidity of motion of the molecules, to which we ascribe the phenomena of heat, increases with the temperature.

72. **Dialysis.**—If, in the experiment of the last section, we should interpose a porous partition of earthenware or parchment between the salt solution and the pure water, the result would be the same; the salt would pass through the partition into the water. If, however, we use on one side of the partition a colloidal substance like gelatine or albumen, we shall find that almost none of this body passes through it into the water. Easily crystallizable bodies pass through the membranes with ease, while those which do not crystallize—called colloids—pass with great difficulty. This property of bodies is made use of to separate the one class from the other. The process of the passage of liquids through porous membranes is called **osmosis**. The **dialyser** is an apparatus consisting of a shallow vessel provided with a bottom of parchment or some porous membrane, into which the solution to be dialysed is placed, and the vessel is then floated upon pure water in a larger vessel. (See Fig. 15.) The volume of the water should be 8 or 10 times that of the solution to be dialysed. In the course of two or three days the crystalline bodies in the solution will be found in the water of the outer vessel, and the colloid bodies will be in the dialyser.

FIG. 15.



Dialysis is employed to prepare a pure colloid material by dialysing from it all crystallizable salts. A complex mixture, like the contents of a stomach, is submitted to dialysis for the purpose of separating from it the crystalloids which it may contain, so as to get them in a pure watery solution for analysis. Recently a series of "**dialysates**" of the alkaloids have found their way into the market. These are the dialysed infusions of the crude drug.

CRYSTALLOGRAPHY.

73. **Formation of Crystals.**—When substances change to the solid form, from the melted state or from a solution, many of them assume some regular geometrical form, called a crystal. The process is called **crystallization**. The same substance

[illegible] when formed in the [illegible] as high and low [illegible] different crystalline [illegible] **amorphous.** Different salts [illegible] the structure of [illegible] Thus Na Cl, K Cl, Na Br, K Br, [illegible] while K_2SO_4, KNO_3, and [illegible] form of the crystal, therefore, [illegible] index of the molecular structure of the body. There are some substances, like gelatin, albumin, starch, etc., which cannot be made to assume the crystalline form. Such bodies are called **colloids,** while those which crystallize readily are called **crystalloids.** In order that crystals may form, it is necessary that the molecules shall be free to move, i. e., cohesion must be overcome to such an extent that it shall not prevent free movement of the molecules. This condition prevails in solutions, in the melted state, or in the gaseous state. When we evaporate off the solvent, in case of solutions, we may obtain the crystals with ease. The more slowly the evaporation takes place, the larger and more perfect are the crystals obtained. As a rule, bodies dissolve more readily and in larger quantities in hot than in cold water. A hot saturated solution of any crystallizable salt deposits the excess, on cooling, in the form of crystals. A liquid which is depositing crystals will form more readily when foreign bodies—as sticks, strings, [illegible] suspended in it. Advantage is taken of this fact to prepare [illegible] ornaments in the shape of grass, leaves, etc., [illegible] with alum crystals, which may be colored with aniline [illegible] by previously coloring the solution.

[illegible] this way, upon strips of [illegible]

[illegible] assume the form of [illegible] for example. [illegible] a solution con- [illegible] of solubility, the [illegible] may thus be separated from [illegible]

[illegible] salt from sea [illegible] the mag- [illegible]

[illegible]

are nearly free from impurities. We therefore take advantage of this method to purify substances.

74. **Water of Crystallization.**—Most substances, when they separate from a solution, take with them a certain definite amount of water as a necessary part of the crystal. This water is known as **water of crystallization.** The crystals of a given substance, when deposited at the same temperature, always contain the same amount of water. Thus, the crystals of copper sulphate contain five molecules of water for one of the salt, and the formula of the crystal is written thus :—

Cu So_4. $5H_2O$ or Cu SO_4. 5Aq.

Ferrous sulphate	crystals	have the	formula	Fe SO_4. $7H_2O$
Sodium carbonate	"	"	"	Na_2CO_3. $10H_2O$
Alum	"	"	"	$K_2Al_2(SO_4)_4$. $24H_2O$

A few salts have different amounts of water of crystallization when separated at different temperatures. Thus crystals of manganous sulphate have the formula—

Mn SO_4. $7H_2O$	when	crystallized	below	6°C.	(42.8°F.)
Mn SO_4. $5H_2O$	"	"	between	7° and 20°C.	(44° and 68°F.)
Mn SO_4. $4H_2O$	"	"	"	20° and 30°C.	(68° and 86°F.)

The crystalline forms in these three cases are entirely different, showing that the molecules of water are necessary to the form of the crystal. The water, in these cases, is held by a feeble force, and may generally be driven off by exposing the crystal to a temperature of 100°C (212°F) in a dry atmosphere, when the crystals fall to powder. In some cases the crystals lose their water at ordinary temperatures and crumble to a powder. It is then said to **effloresce.** On the other hand, some dry substances, when exposed to the air, absorb water. They are then said to **deliquesce**; such a body is said to be **deliquescent** or **hygroscopic.**

Silver nitrate, (Ag NO_3), and a few other salts, crystallize without water of crystallization.

75. **Forms of Crystals.**—A great variety of crystalline forms are met with, but for convenience of classification all may be included in six systems. These systems are based upon the number, length, and inclination of certain imaginary lines called axes, passing through the centre of the crystal and connecting opposite parallel sides, or opposite angles.

The First, or Isometric System.—In this system the axes are three in number, arranged at right angles, and of equal length. The simplest form is the cube; the axes in this case join the centres of the opposite sides. Another common form is the octahedron in which the axes join the opposite angles or corners. See Figs. 17 and 18.

FIG. 16.

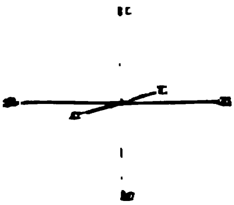
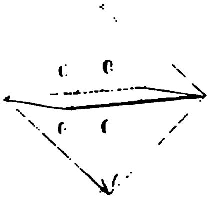
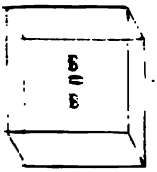


FIG. 17.



REGULAR OCTAHEDRON.

FIG. 18.



CUBE.

NaCl, KCl, KBr, KI, Galena and calcium oxalate crystallize in this system.

76. The Second, Tetragonal, or Dimetric System.—In this system the axes, three in number, are at right angles, as in the last, but one is longer or shorter than the other two. The simplest form is the right square prism; i. e., a prism having a square base and varying in height from a flat plate to a tall column.

FIG. 19.

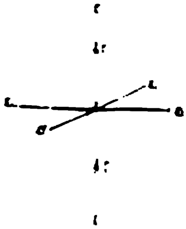
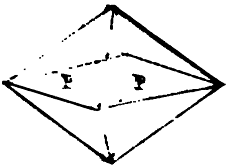


FIG. 20.

FIG. 21.



OCTAHEDRON.

77. The Third, Trimetric, or Orthorhombic System.—In this system the three axes are at right angles, but of unequal lengths. The most characteristic forms of this system are the right rhombic prism and rhombic octahedron. Figs. 22 and 23.

The sulphates of magnesium, zinc, lead, and barium crystallize in this system.

78. The Fourth, or Monoclinic System.—In this system, of the three unequal axes, one only is at right angles to the plane

FIG. 22.

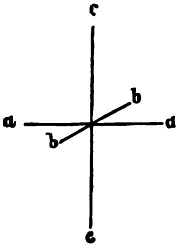


FIG. 23.

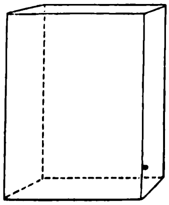
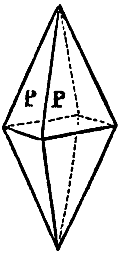


FIG. 24.



of the other two. The simplest form in this system is the oblique rhombic prism. (Fig. 27.) Borax, sulphates of sodium, calcium, and iron crystallize in this system.

FIG. 25.

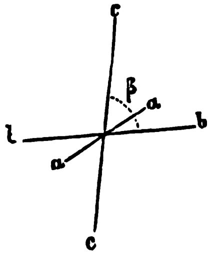


FIG. 26.

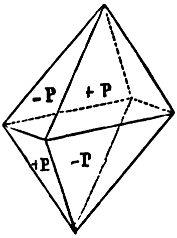
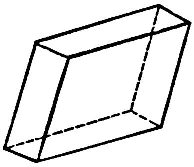


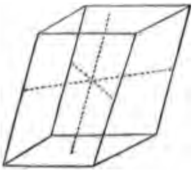
FIG. 27.



OBLIQUE RHOMBIC PRISM.

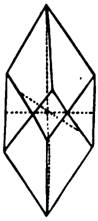
79. The Fifth, or Triclinic System.—In this system the three axes are all unequal, and all inclined to each other. The crystals of this system are usually very complicated, and seemingly irregular. (See Figs. 28, 29 and 30.) Boracic acid, and copper sulphate are examples.

FIG. 28.



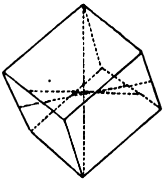
RHOMBOIDAL PRISM.

FIG. 29.



ACUTE RHOMBOHEDRON.

FIG. 30.



OBTUSE RHOMBOHEDRON.

80. **The Sixth, or Hexagonal System.**—In this system there are four axes. Three of them are arranged in one plane, making an angle of 60° with each other, and the fourth, longer or shorter, is set at right angles to the plane of the other three. The prism of this system has eight faces; the bases are regular hexagons, and the other six are parallelograms. The prism of this system is frequently modified by the dodecahedron, as in Fig. 32. Snow crystals, ice, Iceland spar, and quartz, are examples of bodies which crystallize in this system.

FIG. 31.

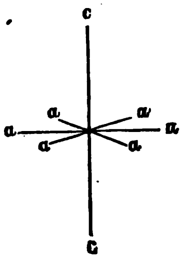
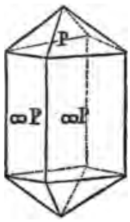
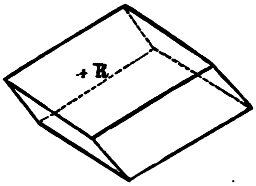


FIG. 32.



HEXAGONAL PRISM.
Modified by Dodecahedron.

FIG. 33.



RHOMBOHEDRON.

PART II.

THEORETICAL CHEMISTRY.

81. **Molecules.**—A **molecule** has been defined as a collection of atoms held together by chemism or affinity, in such a way as to neutralize their tendency to combine with outside atoms.

Hence, a molecule may be defined as the smallest portion of matter that can remain in the free or uncombined state. When the atoms composing a molecule are of the same kind it is said to be **elemental** or **simple**; when of different kinds it is called a **compound molecule.**

When, by chemical means, we cause a re-arrangement of the atoms of compound molecules, we may obtain two or more kinds of elemental molecules; but with elemental molecules we only obtain the same kind. We may illustrate this by the following formulæ:—

Let **ab** and **ab** represent two compound molecules. By a re-arrangement we may have **aa** and **bb**. If we take **aa** and **aa** we will not be able to obtain anything else but **aa** and **aa.**

If we take the molecules represented by HOH and HOH, and cause the re-arrangement by a strong electric current, we shall have **HH, HH** and **OO**, or two kinds of molecules entirely different from the original molecules and from each other.

If, on the other hand, a current of electricity be caused to pass through either HH + HH, or OO + OO, we will only obtain HH and HH, or OO and OO. By this and other methods known to chemists, about 70 elemental molecules or different kinds of atoms have been discovered. By a **chemical element,** then, we mean a substance that has never been found to contain more than one kind of atoms; and a compound body is one that has been shown to contain more than one kind of matter or atoms. These 70 different molecules or atoms have each received a separate name; the name of the molecule and that of the corresponding element being the same. These names are given in the first column of the following table:—

THE CHEMICAL ELEMENTS.

QUANTIVALENCE.	SPECIFIC GRAVITY.	ATOMIC WEIGHT.	REVISED AT. WEIGHT.	ELECTRICAL STATE.
[illegible]	2.5	27.0	[illegible]	—
[illegible]	6.7	120	[illegible]	—
[illegible]	5.8 gas 150	75 L=13.44 ÷	74.918	—
[illegible]	4	137	136.763	+
[illegible]	9.8	210	207.523	—
[illegible]	2.63	11	10.941	—
[illegible]	3.2 D*=80	80 L=7.17	79.768	—
[illegible]	8.6 D=56	112	111.835	—
	...	133	132.583	—
	1.57	40	39.990	+
	D. 3.3 G. 2.3	12	11.973	—
	[illegible]	141.2	140.424	+
	[illegible]	35.5	35.370	—
[illegible]	[illegible]	52.5	52.009	—
	[illegible]	59	58.887	+
	...	94	93.812	—
	[illegible]	63.4	63.173	+
	[illegible]	154	...	—
	[illegible]	144.7	144.573	—
		166.5	165.891	—
		19	18.984	—
		69.9	68.854	+
		14	...	
		196.2	196.155	+
	[illegible]	[illegible]	1.000	+
	[illegible]	113.4	113.398	+
	[illegible]	[illegible]	126.557	—
	[illegible]	[illegible]	...	+
	[illegible]	56	55.913	+
	[illegible]	[illegible]	206.471	—
		[illegible]	138.526	—
		[illegible]	7.007	—
		24	23.959	+

1 Weight [illegible] of vapor or gas.

NAME.	SYMBOL.	QUANTIVALENCE.	SPECIFIC GRAVITY.	ATOMIC WEIGHT.	REVISED ATOMIC WEIGHT.	ELECTRICAL STATE.
34 Manganesium.	Mn.	II, IV, $(Mn_2)^{VI}$	8	55	53.906	+
35 Mercury (Hydrargyrum).	Hg.	$(Hg_2)^{II}$ II.	13.6 (D=100)	200 (L=8.96)	199.712	+
36 Molybdenum	Mo.	*II*, *IV*, VI.	8.6	96	95.527	—
37 Nickel.	Ni.	II, *IV*, $(Ni_2)^{VI}$	8.8	58	57.928	+
38 Nitrogen.	N.	*I*, III, *V*.	14	14	14.021	—
39 Osmium.	Os.	*II*, IV, *VI*.	21.4	199	198.494	+
40 Oxygen.	O.	II.	16	16 (L=1.43)	15.963	—
41 Palladium.	Pd.	*II*, IV.	11.6	106	105.737	+
42 Phosphorus.	P.	III, *V*.	2.2 1.83 D =62	31	30.958	—
43 Platinum.	Pt.	*II*, IV.	21.5	197	196.700	+
44 Potassium (Kalium).	K.	I, *III*, V.	.86 D=39.1	39.1 (L =3.5)	39.019	+
45 Rhodium.	Ro.	*II*, IV, *VI*.	11	104.4	104.055	+
46 Rubidium.	Rb.	I.	...	85.50	85.251	+
47 Ruthenium.	Ru.	*II*, IV, *VI*.	11.4	104.4	104.217	+
48 Selenium.	Se.	II, *IV*, *VI*.	4.8	79	78.797	—
49 Silicon.	Si.	*II*, IV.	2.6	28	28.195	—
50 Silver (Argentum).	Ag.	I, *III*.	10.5	108	107.675	+
51 Sodium (Natrium).	Na.	I, *III*.	.97 D =23	23 (L=2.06)	22.998	+
52 Strontium.	Sr.	II, *IV*.	2.54	87.5	87.374	+
53 Sulphur.	S.	II, *IV*, *VI*.	2 D =32	32 (L=2.86)	31.984	—
54 Tantalum.	Ta.	V.	10.78	182	182.144	—
55 Tellurium.	Te.	II, *IV*, *VI*.	6.6	128	127.960	—
56 Thallium.	Tl.	I, *III*.	11.8	204	203.715	+
57 Terbium.	Ter.	...	7.9	148.5	...	
58 Thorium.	Th.	*II*, IV.	7.9	231	231.500	+
59 Tin (Stannum).	Sn.	*II*, IV.	7.3	118	117.698	+
60 Titanium.	Ti.	*II*, IV.	...	50	49.846	—
61 Tungsten (Wolfram).	Wo.	*IV*, VI.	17.6	184	183.610	—
62 Uranium.	Ur.	II, IV, $(Ur_2)^{VI}$	18.4	240	239.482	+
63 Vanadium.	V.	III, V, $(V_2)^{VI}$ $(V_2)^{VIII}$	...	51.3	51.256	—
64 Yttrium.	Yt.	II.	4.8	92.5	89.816	+
65 Zinc.	Zn.	II.	7 D =32.5	65	64.904	+
66 Zirconium.	Zr.	*II*, IV.	4.1	89.5	89.367	+

The following named elements have been omitted from the above table, their identity not being yet thoroughly established :—

NAME.	SYMBOL.	QUANTIVALENCE.	ATOMIC WEIGHT.
67 Decipium.	Dp.	*II.*	160
68 Holmium.	Ho.		162
69 Ilmenium.	Il.	*V.*	105
70 Lavorsium.	Lv.	?	?
71 Mosandrium.	Ms.	?	?
72 Neptunium.	Np.	*V.*	118
73 Philipium.	Pp.	*II,*	74
74 Scandium.	Sc.	*II.*	44
75 Thulium.	Tm.		170
76 Ytterbium.	Yb.	*II.*	173?

83. **Molecular Weights.**—Molecules, whether elemental or compound, must have a definite size and weight. The absolute weight of molecules and atoms is of no practical value to the chemist, but the comparative weights of molecules we shall find to be of vast importance. In weighing molecules we use the lightest known atom as the unit of weight. This atom is that of hydrogen. The relative weights of molecules have all been measured, and in expressing their weights our numbers express how many times heavier the molecule is than the hydrogen atom. Thus, the molecular weight of oxygen is 32. That is, the molecule of oxygen weighs as much as 32 atoms of hydrogen.

Molecular weight, then, is the weight of a molecule as compared with the weight of the hydrogen atom.

84. **Avogadro's or Ampere's Law.**—This law was first enunciated by Avogadro, an Italian physicist, in 1811, and was reproduced by Ampére, a French physicist, in 1814. The law has already been stated (see chapter on gaseous state) as follows :— Equal volumes of all true gases, when at the same temperature and under the same pressure, contain the same number of molecules. That is to say, a litre of any given gas, under the same conditions of temperature and pressure, always contains the same number of molecules, whatever the nature or composition of the molecules composing the gas. As a natural conclusion from this law, we have the following : **First.**—Gaseous molecules always occupy the same space , *i. e.*, the molecule, together h the intervening space, always occupies the same volume.

Second.—Since the same volume contains the same number of molecules, it follows that the weights of equal volumes of any two gases (under like conditions as above), will be the weights of the same number of molecules. Hence, the two weights will stand in the same proportion to each other as the weights of their molecules.

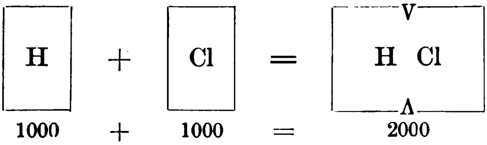
Thus, suppose equal volumes of hydrogen and oxygen gases, large enough to contain 10,000 molecules each; the weights of these gases will be in the one case 10,000 times the weight of one molecule of hydrogen, and in the other 10,000 times the weight of one molecule of oxygen. and these numbers must be to each other as the weight of 1 molecule of hydrogen to 1 molecule of oxygen. This law is the basis of many of our modern chemical notions. It is to chemistry what Newton's law of gravitation was to the science of astronomy. This is not the place to enter into the discussion of the proofs of this law. Suffice it to say, that it rests on about as strong proof as any other law of physics or chemistry, or as the law of gravitation itself. As the present science of astronomy is the living proof of the one, so the present science of chemistry is one combined proof of the other.

85. **Number of Atoms in Elemental Molecules.**—We may determine the number of atoms in many elemental molecules by a simple application of Avogadro's law to well known experiments.

This demonstration may be illustrated by reference to the behavior of the two gases, chlorine and hydrogen.

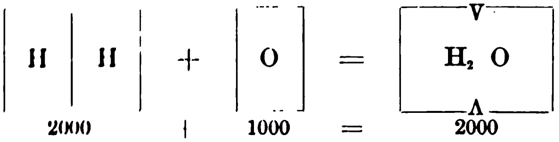
Into a glass tube inverted over mercury, put equal volumes of the two gases, and allow the apparatus to stand in diffuse light. After some hours, the greenish color of the mixture will have entirely disappeared, the gases having combined to form the colorless hydrochloric acid. The mercury stands at the same height in the tube as at the beginning of the experiment. The volume of the hydrochloric acid is therefore just equal to the volume of both constituents.

Let the following diagram represent this combination:—



Since the volume of the hydrogen and that of the chlorine are equal, it is clear, from the law, that the number of molecules of each will be the same; also, the number of molecules of the hydrochloric acid gas must be equal to the sum of the molecules of the two gases used. Suppose, for illustration, that the volume of hydrogen taken contained 1000 molecules; then there will be 1000 molecules of chlorine, and 2000 molecules of hydrochloric acid, the volume of this gas being twice that of hydrogen. Each of the 2000 molecules of hydrochloric acid must contain an atom of hydrogen and one of chlorine; hence, 2000 atoms of each of these elements have been derived from 1000 molecules of the same, and therefore each molecule must have furnished two atoms.

Experiments quite as decisive have been made to show that in water, the 2 volumes of hydrogen and 1 volume of oxygen produce only 2 volumes of water-vapor. Thus:—



By the same reasoning we may show that each molecule of oxygen contains two atoms; also, that each molecule of water must contain *two* atoms of hydrogen to one of oxygen, instead of *one*, as formerly taught.

Most of the molecules of elementary substances contain two atoms; or, in other words, are diatomic. Mercury, cadmium, zinc, and barium, however, contain but one atom in each molecule. Oxygen, as **ozone**, contains three atoms. Phosphorus and arsenic contain four atoms. Sulphur at lower temperatures contains six, while at higher temperatures it contains only two.

Some molecules which at ordinary temperatures contain two atoms in their molecules, break up or **dissociate**, as it is termed, at very high temperatures. Thus chlorine, bromine, and iodine, which at a moderate heat are diatomic, at a very high temperature, 800 C., begin to dissociate, and become monatomic at the strongest furnace heat. Many compounds do the same thing, (Hg_2 Cl_2, N_2O_2, N_2O_4). The law of Ampere is, therefore, true only at moderate temperatures.

80. **Molecular Weight, Determined by the Law of Avogadro.** It has just been shown that the molecule of hydrogen is composed of two atoms. We have already defined molecular weight, as the weight of a molecule in units of hydrogen atoms. It may also be defined as the sum of the weights of its constituent atoms. The molecular weight of hydrogen is therefore 1 + 1, or 2. Suppose, for example, we weigh equal volumes of hydrogen and oxygen gases under like conditions of temperature and pressure, and find the weights to be respectively 1 and 16 grammes. It follows that the molecules of these gases are to each other as 1 to 16, for, according to the law, each contains the same number of molecules. As the molecular weight of hydrogen has been shown to be 2, it follows that the molecular weight of oxygen must be 32, for 32 bears the

same ratio to 2, that 16 does to 1. 1 : 16 :: 2 : 32. From this it will be seen that the molecular weight of oxygen is twice its density, or its specific gravity as compared with hydrogen. The same reasoning applies to all other gases, whether elemental or compound, and we may state this fact as follows: **The molecular weight of any body is twice its density in the state of gas.** The converse of this statement is also true, viz.: **the density of a gas is one-half its molecular weight.**

PROPERTIES OF ATOMS.

87. **Definition.**—An atom is the smallest conceivable portion of matter. It is the smallest portion of an element that can enter into the formation of a molecule, or take part in a chemical reaction.

The student should strive to clearly comprehend the difference between a molecule and an atom. The former is a collection of the latter held together by an attraction called chemism or chemical affinity. Molecules are destructible; they may be broken up and their properties destroyed. The atom, on the other hand, is an indestructible solid particle, whose properties, so far as we know, are never destroyed.

As to the properties of molecules, we conceive that they have size, weight, form, color, etc.

88. **Atomic Weight.**—Atoms differ from one another in their weight, and the quality and the quantity of their combining power. Further than this we have no certain knowledge of their properties.

The weight of an atom of any given element is always the same. In weighing atoms we do not take the absolute weight, but the relative weight, using the hydrogen atom as the unit. The **atomic weight** of any element expresses the number of times its atom is heavier than the atom of hydrogen. The atomic weight of oxygen is 16; *i. e.*, the atom weighs 16 times as much as the atom of hydrogen. The atomic weights of the elements will be found in the table, Art. 82.

89. **Quality of Combining Power.—Polarity of the Atoms.**—We have seen (Art. 64) that when a current of electricity is sent through a solution of a metallic salt, the metal collects about the negative electrode, while the non-metallic part of the salt appears at the positive.

Most metallic compounds are capable of decomposition by electrolysis, and the molecule seems to be divided into two parts; one of which is attracted to the positive and the other to the negative electrode. We have seen (Art. 44) that two bodies, similarly polarized, repel each other, while bodies oppositely polarized attract each other. We conclude, therefore, that those atoms which are attracted by the negative pole of the galvanic current are positively polarized, while the others are negatively polarized. This is what is meant by the difference in the quality of combining power. Positive and negative, as applied to the polarity of atoms, is not absolute but relative; and the polarity of an atom may be changed, by the inducing action of another atom, from positive to negative, or vice versa. As a general rule, the metallic atoms are positive, and the non-metallic negative.

When several atoms are brought into contact with each other, those having similar polarity repel one another, as do other bodies that are of like polarity, while those having differing polarity attract one another; hence, union or combination can only take place between atoms that are unlike in their electric quality.

In the following table the elements are so placed that each one is electro-negative to those below, and electro-positive to those above it.

ELECTRO-CHEMICAL SERIES.

Negative End —.
Oxygen.
Sulphur.
Nitrogen.
Fluorine.
Chlorine.
Bromine.
Iodine.
Selenium.
Phosphorus.
Arsenic.
Chromium.
Vanadium.
Molybdenum.
Tungsten.
Boron.
Carbon.
Antimony.
Tellurium.
Tantalum.
Columbium.
Titanium.
Positive End +.

Negative End —.
Silicon.
Hydrogen.
Gold.
Osmium.
Iridium.
Platinum.
Rhodium.
Ruthenium.
Palladium.
Mercury.
Silver.
Copper.
Uranium.
Bismuth.
Tin.
Indium.
Lead.
Cadmium.
Thallium.
Cobalt.
Nickel.
Positive End +.

Negative End —.
Iron.
Zinc.
Manganese.
Lanthanum.
Didymium.
Cerium.
Thorium.
Zirconium.
Aluminum.
Erbium.
Yttrium.
Glucinum.
Magnesium.
Calcium.
Strontium.
Barium.
Lithium.
Sodium.
Potassium.
Rubidium.
Cæsium.
Positive End +.

90. **Quantity of Combining Power.**—By an analysis of a large number of compounds of hydrogen with other elements, it has been found that while chlorine combines with it in the pro-

portion of its atomic weight, *i.e.*, 1 part of hydrogen to 35.5 parts of Cl,

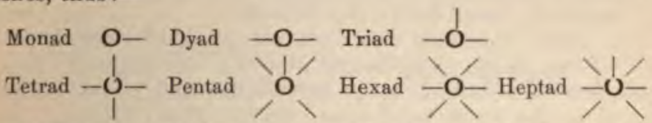
Oxygen	combines with	2	parts of H to	16 of O,
Nitrogen	"	" 3	" of H to	14 of N,
Carbon	"	" 4	" " "	12 " C,

and so on; so that the power of the atoms to attract and combine with hydrogen are not alike. This is expressed by saying that the equivalence, or quantivalence, of the atom in question is 1, 2, 3, 4, 5, 6 or 7, according as it will attach to itself, be exchanged for, or take the place in a molecule, of 1, 2, 3, 4, 5, 6, 7, atoms of hydrogen, or their equivalent.

The chemical equivalent of an atom, is an atom which can take its place in a molecule. Atoms are divided into monads, dyads, triads, tetrads, pentads, hexads, or heptads, according as they can fix 1, 2, 3, 4, 5, 6 or 7 atoms of hydrogen, or their equivalent.

A monad	is	equivalent	to a monad.	
" dyad	"	"	to 2 monads.	
" triad	"	"	to 3 "	or 1 monad and 1 dyad.
" tetrad	"	"	to 4 "	2 dyads, or 1 monad and 1 triad.
A pentad	"	"	to 5 "	a tetrad and 1 monad, a triad and 2 monads, or 2 dyads and 1 monad.

The equivalence of an atom is often indicated to the eye by dashes, thus:—



It will be seen that the hydrogen atom is the unit of comparison in combining powers or equivalences, and the dashes represent the number of **bonds** or **points of attraction**, or **poles** of the **atomic magnet.** The equivalence of an atom may also be expressed by a Roman numeral placed above and to the right, thus:—H^{I}, O^{II}, Cl^{III}, C^{IV}, etc. Univalent, bivalent, trivalent, etc., are adjectives sometimes used as synonymous with monad, dyad, triad, etc.

CHEMICAL NOTATION.

91. **Symbols and Formulæ.**—In representing atoms and molecules to the eye, we make use of a series of symbols derived from the names of the elements themselves. This is usually the initial letter of the English or Latin name, or in case two or

more names begin with the same letter, the initial with some other characteristic letter. Thus, on reference to the table in Art. 82, we have B, Ba, Bi, Br, representing the atoms of boron, barium, bismuth and bromine respectively. In this book, a symbol is never used to represent the element in general, but a symbol always represents an atom, with all its properties, and nothing else.

Formulæ.—A formula is the sign of a molecule. It therefore, represents a definite weight—the molecular weight; and in the case of gases, always the same volume. Formulæ are made up of symbols, as a molecule is made up of atoms; and the atoms composing a molecule are all represented by symbols in the formula. Thus: HCl is a formula representing a molecule containing one atom or 1 part of hydrogen, and one atom or 35.5 parts of chlorine.

In writing formulæ, we write the symbols composing the molecule in juxtaposition, beginning with the more electro-positive. Thus: $KOCl$, HBr, etc.

Multiplication of Molecules and Atoms.—When we wish to represent more than one atom, we use a small numeral at the right hand lower corner; thus, O_2, represents 2 atoms of oxygen, or, since the molecule of oxygen contains two atoms, this also represents the molecule. As_4 represents 4 atoms of arsenic, or since the molecule of arsenic contains four atoms, it is also the formula of a molecule of arsenic. When we wish to represent more than one molecule of a substance, we use full-sized numerals placed before the formula.

Thus, $2H_2O$ represents two molecules, each of which is composed of two atoms of hydrogen and one of oxygen. Or, we may inclose the formula in a parenthesis, and place a small numeral at the right hand lower corner, thus: $(H_2O)_2$. Examples:—

H_2SO_4 represents 1 molecule, containing 2 atoms of hydrogen, 1 of sulphur, and 4 of oxygen.

$5H_2SO_4$ represents 5 molecules of the same substance.

$3NH_4NO_3$ represents 3 molecules, containing in each molecule two atoms of nitrogen, 4 atoms of hydrogen, and 3 atoms of oxygen; 27 atoms in all.

$2K_2Al_2(SO_4)_4$ represents two molecules, containing in each molecule 2 atoms of potassium, 2 atoms of aluminium, 4 atoms of sulphur, and 16 atoms of oxygen; 48 atoms in all.

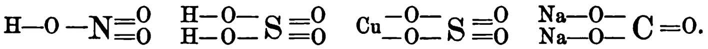
As the symbols always represent the atomic weights, we may reduce any formula to figures, or find its molecular weight, by adding together the weights of the symbols composing it. Let it be desired to find the molecular weight of H_2SO_4. By reference to the table in Art. 82, it will be seen that $H_2 = 2$, $S = 32$

and $O_4 = 4 \times 16 = 64$. By adding together these three numbers we obtain 98, the weight of the molecule. An **empirical formula** is one which merely gives the kind and number of the atoms composing a molecule. A **rational** or graphic **formula** aims to show the arrangement of the atoms in the molecule, with relation to one another.

Examples of empirical formulæ:—

HNO_3, H_2SO_4, $CuSO_4$ Na_2CO_3.

Rational formulæ for the same:—

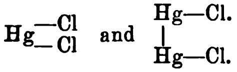


$$H-O-N\begin{matrix}=O\\=O\end{matrix}\quad \begin{matrix}H-O-\\H-O-\end{matrix}S\begin{matrix}=O\\=O\end{matrix}\quad Cu\begin{matrix}-O-\\-O-\end{matrix}S\begin{matrix}=O\\=O\end{matrix}\quad \begin{matrix}Na-O-\\Na-O-\end{matrix}C=O.$$

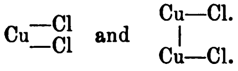
Rational formulæ are useful in giving us a more definite conception of the relations of the atoms to one another in the molecule. They have served as the guides in some of the most important chemical discoveries of the present century; such as the discovery of the process of manufacturing artificial madder and indigo from coal tar products, by synthesis.

92. **Variation in Quantivalence.**—By graphic formulæ we are able to explain a fact that is always a matter of difficulty to the student, viz.: the variation of equivalence of atoms.

There are two well known series of salts of mercury and copper, in which there is no real variation; but, owing to the uniting of two atoms of the metals, each loses an available bond or point of attraction. The following formulæ will render this clear:—



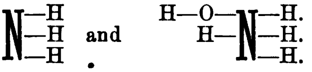
$$Hg\begin{matrix}-Cl\\-Cl\end{matrix}\quad\text{and}\quad\begin{matrix}Hg-Cl.\\|\\Hg-Cl.\end{matrix}$$



$$Cu\begin{matrix}-Cl\\-Cl\end{matrix}\quad\text{and}\quad\begin{matrix}Cu-Cl.\\|\\Cu-Cl.\end{matrix}$$

In other cases, and under certain well known conditions which we can control in the laboratory, the atom which has previously existed as a dyad, suddenly becomes a tetrad, or a triad becomes a pentad, and so on.

These changes are always extremely puzzling to the student, and we shall dwell a little upon them. When ammonia gas for instance, (NH_3), is absorbed by water, it combines with a molecule of the water and becomes NH_4OH. If we represent the two molecules graphically, we have,



$$N\begin{matrix}-H\\-H\\-H\end{matrix}\quad\text{and}\quad\begin{matrix}H-O-\\H-\\\end{matrix}N\begin{matrix}-H.\\-H.\\-H.\end{matrix}$$

As will be seen, two new points of attraction have made their appearance upon the nitrogen atom. A large number of such cases are known, and the explanation is as follows:—

The full quantivalence, or atomicity of nitrogen, is pentad. In the compound H_3N, for some unknown reason, two poles of the atomic magnet neutralize each other, and so the combining power of the atom is lessened by two. This increase or diminution of combining power always takes place in pairs, so that a dyad may become a tetrad, but not a triad. A monad may become a triad or a pentad but never a dyad or tetrad.

93. **Other Signs used in Writing.**—A **plus sign** between two formulæ indicates that the substances, whose molecules they represent, are brought together.

The **minus sign** indicates that the molecule following it is abstracted from the preceding one. The **sign of equality** is used to indicate that what follows, is the result of some change that has taken place. $HCl + AgNO_3 = HNO_3 + AgCl$ shows that the molecules represented by the first two formulæ have been brought together, and that a change has taken place resulting in the formation of the last two.

94. **Compound Radicals.**—A radical or root of a series of compounds, is a characteristic atom or group running through all of them, like a **root** in language. Thus, the interrogative root runs through all that class of words, as **who, which, when, what, etc.** So, in chemical compounds we have a large number of potassium compounds, in which the atom K appears as the characteristic atom: As KNO_3, $KClO_3$, K_2CO_3, K_2SO_4 and KCl. It is therefore called the root or radical of these compounds. A single atom, which forms a series of characteristic compounds, is called a **simple radical.**

Sometimes, instead of being a single atom, it is a group of atoms that is found to be the characteristic of a series of compounds. Thus, we have: $(NH_4)NO_3$, $(NH_4)Cl$, $(NH_4)NO_2$, $(NH_4)_2S$, etc., in which the characteristic radical is a group of atoms, or is a **compound radical.**

A compound radical may be regarded as a group of atoms which behaves like a simple radical, or single atom. Like the single atom it only exists in combination with another atom or group of atoms; for its bonds or points of attraction are not satisfied unless it be in combination. Compound radicals, like atoms, may be positive or negative. Each compound radical has a definite equivalence like the atoms. Some of them have received arbitrary names which do not express their composition, and in most cases end in yl. Thus, $(PO)'''$, phosphoryl,

$(H\text{-}O\text{-})'$ hydroxyl, $(CO)''$ carbonyl, $(CH_3)'$ methyl, $(C_2H_5)'$ ethyl, $(H_4N)'$ ammonium, $(CN)'$ cyanogen, $(NH_2)'$ amidogen. The last three are exceptions to the rule as to the ending.

In writing the formula of these compound radicals, they may be regarded for the time as atoms of a compound nature. If we wish to represent that several similar compound radicals enter into the same molecule, we inclose the formula in a parenthesis, and use the numerals as with atoms, thus, $(NH_4)_2CO_3$, $Fe_2(OH)_6$.

In the following table will be found the more important elements arranged according to both quality and quantity of combining power. The elements at the top of the table are negative to all below them; and those at the bottom are positive to all above. They are also divided into monad, dyad, triad, etc., some appearing in two or even three columns, because of their change of quantivalence. In the second table will be found the more common compound radicals, with the names of the classes of compounds they form, arranged, as far as possible, in the same order as the elements. In regard to their electrical order, much less certainty exists than with the elements.

QUANTIVALENCE OF THE ELEMENTS.

MONADS.	DYADS.	TRIADS.	TETRADS.	PENTADS.	HEXAD.
	Oxygen...... O				Sulphur.... S
	Sulphur..... S		Sulphur........ S	Nitrogen ... N	
Fluorine ... F		Nitrogen. N			
Chlorine ... Cl					
Bromine ... Br		Phosphorus P		Phosphorus P	
Iodine I		Arsenic... As		Arsenic....... As	Chromium Cr
		Boron..... B			Manganese Mn
		Antimony Sb	Carbon......... C		
			Silicon......... Si		
Hydrogen.. H					
		Gold ...... Au	Platinum...... Pt		
		Bismuth. Bi	Tin.............. Sn		
	Mercury ... Hg		Iron .... Fe		
Silver........ Ag	Copper...... Cu		Chromium.... Cr		
	Lead......... Pb		Manganese... Mn		
	Cadmium .. Cd				
	Cobalt....... Co				
	Nickel ...... Ni				
	Iron .... Fe				
	Chromium. Cr				
	Manganese Mn				
	Zinc.... .... Zn		Aluminum.... Al		
	Magnesium Mg				
	Calcium .... Ca				
	Strontium.. Sr				
	Barium ..... Ba				
Sodium..... Na					
Potassium. K					

COMPOUND RADICALS WITH NAMES OF COMPOUNDS THEY FORM.

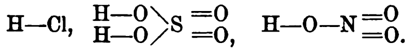
MONAD.	DYAD.	TRIADS.	TETRADS.
NO_3=Nitrates	SO_4=Sulphates	PO_4=Phosphates	$Si\,O_4$=Silicates
$Cl\,O_3$ Chlorates	$Cr\,O_4$=Chromates	BO_3=Borates	P_2O_7=Pyrophosphates
NO_2 Nitrites	$Cr_2\,O_7$=Bichromates	$As\,O_3$=Arsenites	$Fe\,Cy_6$=Ferrocyanides
$Cl\,O$=Hypochlorites	SO_3=Sulphites	$As\,O_4$=Arsenates	
PH_2O_2=Hypophosphites	CO_3=Carbonates		
CN Cyanides	C_2O_4=Oxalates		HEXADS.
$C_2H_3O_2$ Acetates	$C_4H_4O_6$=Tartrates	$C_6H_5O_7$=Citrates	
$C_7H_5O_2$ Benzoates	$C_4H_4O_5$=Malates		
C_6H_5O Carbolates	$C_7H_4O_3$=Salicylates		Fe_2Cy_{12}=Ferricyanides
$C_3H_5O_3$= Lactates			Al_2=Aluminic Salts
NH_4 Ammonium			Cr_2=Chromic "
CH_3 Methyl	Hg_2=Mercurous Salts		Mn_2=Manganic "
C_2H_5=Ethyl	Cu_2=Cuprous Salts		Fe_2=Ferric "

COMPOUND MOLECULES.

95. **Compound Molecules Classified.**—The system of nomenclature now in use for naming chemical compounds, is based upon the composition and properties of the bodies in question; and the name of a body is intended to express our idea of its chemical composition. Homogeneous bodies are supposed to be made up of a collection of similar molecules; hence a formula which represents the composition of a single molecule, really represents the composition of the mass. In applying names to compounds, we apply the name to the molecule as well as to the mass.

Compound bodies may be divided into two classes: 1st. Those whose molecule is composed of two kinds of atoms or radicals, called **binary compounds**, and 2d, those whose molecules are composed of three or more kinds of atoms or radicals, called **ternary molecules.** Examples: $NaCl$, KBr, $MgCl_2$, and $(NH_4)Cl$ are examples of binary molecules. $KClO_3$, K_2SO_4, $CaCO_3$, $(NH_4)NO_3$ and $Ba(NO_3)_2$ are examples of ternary molecules.

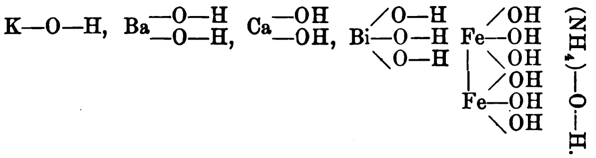
Acids, Bases and Salts.—Ternary molecules are divided into **acids, bases,** and **neutrals** or **salts.** An acid is a substance which usually possesses a sour taste, corrodes the metals with the evolution of hydrogen and the formation of salts, changes blue vegetable colors to reds, and neutralizes the caustic properties of alkalies by forming salts with them. All acids contain hydrogen, which can be replaced by a metal. This hydrogen is united to the remaining portion of the molecule, either directly, as in binary acids, or by a linking atom, usually oxygen, as represented by the following graphic formulæ:



$$\mathrm{H{-}Cl},\quad \left.\begin{array}{l}\mathrm{H{-}O}\\ \mathrm{H{-}O}\end{array}\right\rangle \mathrm{S}\begin{array}{l}=\mathrm{O}\\ =\mathrm{O},\end{array}\quad \mathrm{H{-}O{-}N}\begin{array}{l}=\mathrm{O}\\ =\mathrm{O}.\end{array}$$

The replaceable hydrogen of an acid is called **basic hydrogen**, and the number of such atoms determines the **basicity** of the acid. A dibasic acid, for example, is one containing two atoms of basic hydrogen ; a tri-basic acid three, a tetra-basic acid four, and so on. When the linking atom in these ternary acids is oxygen, the name of **ox-acids** is applied to them. The term **sulpho-acids** is applied to those containing linking sulphur. A **base** has properties which in many respects are opposed to, and neutralize the effects of acids. They restore the vegetable blue colors reddened by acids, they neutralize the sour taste, they react upon acids to form salts, with the elimination of one or more molecules of water. The strong bases have a caustic action upon the tissues, decomposing the fats, with which they form soaps.

A **base** may be defined, as a compound whose molecule is composed of a positive atom, or group of atoms, united by linking oxygen to hydrogen. The positive atom is, in most cases, metallic. As :



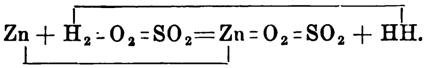
$$\mathrm{K{-}O{-}H},\quad \mathrm{Ba}\begin{array}{l}-\mathrm{O{-}H}\\ -\mathrm{O{-}H},\end{array}\quad \mathrm{Ca}\begin{array}{l}-\mathrm{OH}\\ -\mathrm{OH},\end{array}\quad \mathrm{Bi}\begin{array}{l}/\mathrm{O{-}H}\\ -\mathrm{O{-}H}\\ \backslash\mathrm{O{-}H}\end{array}\quad \begin{array}{l}\mathrm{Fe}\begin{array}{l}/\mathrm{OH}\\ -\mathrm{OH}\\ \backslash\mathrm{OH}\end{array}\\ \quad|\\ \mathrm{Fe}\begin{array}{l}/\mathrm{OH}\\ -\mathrm{OH}\\ \backslash\mathrm{OH}\end{array}\end{array}\quad (\mathrm{NH_4}){-}\mathrm{O}{-}\mathrm{H}.$$

In the last formula we have an example of a compound radical united to H by O. The bases are named hydrates.*

A **salt** molecule is composed of a positive radical united by linking oxygen to a negative radical. The radicals, in this case, as in acids and bases, may be either simple or compound. Thus:

$\mathrm{K{-}O{-}Cl}$, $\mathrm{K{-}O{-}NO_2}$, $\mathrm{Na_2{=}O_2{=}CO}$ $\mathrm{Ba{=}O_2{=}SO_2}$$\mathrm{(NH_4){-}O{-}NO_3}$.

It is evident, also, that a salt may be formed by treating an acid with a metal which replaces the hydrogen of the acid with metallic atoms.



$$\mathrm{Zn} + \mathrm{H_2{=}O_2{=}SO_2} = \mathrm{Zn{=}O_2{=}SO_2} + \mathrm{HH}.$$

It may be regarded, then, as an acid whose replaceable hydrogen atoms have been replaced by positive atoms or radicals. In a dibasic acid, like $\begin{array}{l}H—O\\H—O\end{array}\!\!>S<\!\!\begin{array}{l}O\\O,\end{array}$ it is possible to replace one of the atoms of hydrogen and leave the other undisturbed. We thus have, for example, $\begin{array}{l}H—O\\K—O\end{array}\!\!>S<\!\!\begin{array}{l}O\\O,\end{array}$ which exhibits the properties, and answers to the definition of both a salt and an acid. It has acid properties by virtue of the replaceable hydrogen, and saline properties by virtue of the other chain in which the K has replaced H.

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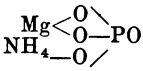
* The above definition applies only to inorganic bases.

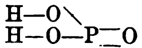
Such a body is called an **acid salt**, while the salts first mentioned, in which all the H atoms have been replaced by positive atoms, are called **normal salts.**

Double salts are formed by replacing a part of the hydrogen of the acid by one positive radical, and a part by another.

$KNaSO_4$ from $HHSO_4$,

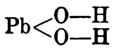


$\begin{array}{l}Mg<\begin{array}{l}O\\O\end{array}\\NH_4—O\end{array}\!\!>PO$ from

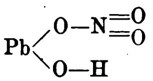


$\begin{array}{l}H—O\\H—O\\H—O\end{array}\!\!>P{=}O$

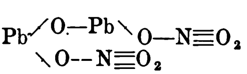
If a base or metallic oxide be treated with sufficient acid to neutralize it, a neutral salt is **usually** formed; but if the base or oxide be much in excess of what the acid would require to expel all its hydrogen, a basic salt will in some cases be formed, according to the following formulæ:—



$Pb<\begin{array}{l}O—H\\O—H\end{array}$



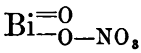
$Pb<\begin{array}{l}O—N{=}O\,{=}O\\O—H\end{array}$ or



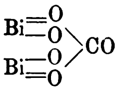
$Pb<\begin{array}{l}O—Pb—O—N{\equiv}O_2\\O—N{\equiv}O_2\end{array}$

i. e., if the base be used, the acid will take the place of a part of its replaceable hydrogen, and leave a part of it; or, a part of the excess of oxide will crowd into the molecule between the negative radical and the positive. Such bodies are called **basic** or **subsalts.**

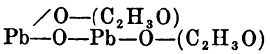
The **subsalts** are seldom of definite chemical composition, often being mixtures of the oxide with the basic or even normal salt Lead and bismuth are two metals especially liable to form basic salts.



$Bi\begin{array}{l}{=}O\\—O—NO_3\end{array}$



$\begin{array}{l}Bi\begin{array}{l}{=}O\\—O\end{array}\\Bi\begin{array}{l}—O\\{=}O\end{array}\end{array}\!\!>CO$



$Pb\begin{array}{l}—O—(C_2H_3O)\\—O—Pb—O—(C_2H_3O)\end{array}$

NOMENCLATURE.

96. **Naming of Chemical Compounds.—Rule.—Give the name of the positive radical first; then the name of the leading negative atom or radical with its termination changed to ide, in binaries, and to ite or ate in ternaries; ite denoting the lower, and ate the higher quantivalence of the negative atom.**

EXAMPLES.

$Na\,Cl$ = Sodium Chloride, — binary.
$Na\,NO_2$ = “ Nitrite, — ternary.
$Na\,NO_3$ = “ Nitrate, — “
$Ba\,Cl_2$ = Barium Chloride, — binary.
$Ca\,Br_2$ = Calcium Bromide, — “
$Ba\,SO_3$ = Barium Sulphite — ternary.
$Ba\,SO_4$ = “ Sulphate, — “

As will be seen on inspection, the equivalence of the negative atom is indicated by the comparative amounts of oxygen which it holds. Compare $BaSO_3$ and $BaSO_4$; also $NaNO_2$ and $NaNO_3$. In compounds like the following—$SnCl_2$, $SnCl_4$, $CuCl_2$, Cu_2Cl_2, $HgCl_2$, Hg_2Cl_2, $FeCl_2$ and Fe_2Cl_6—where more than one quantivalence of the positive atom is known, the endings **ous** and **ic** are used to distinguish them; thus:—

$SnCl_2$ = Stann*ous* Chloride, dyad tin.
$SnCl_4$ = Stann*ic* “ tetrad tin.
$CuCl_2$ = Cupr*ic* “
Cu_2Cl_2 = Cupr*ous* “
Hg_2Cl_2 = Mercur*ous* “
$HgCl_2$ = Mercur*ic* “
$FeCl_2$ = Ferr*ous* “ dyad iron.
Fe_2Cl_6 = Ferr*ic* “ tetrad iron.

Where more than two quantivalences are known, we employ the prefix **per** to denote a quantivalence higher than that denoted by **ic** or **ate**, and **hypo** to denote a lower than that denoted by **ous** or **ite**. They are affixed alike to positive and negative.

EXAMPLES.

N_2O = Nitrous oxide.
N_2O_2 = Nitric “
N_2O_4 = Nitric peroxide or tetroxide.
$Na_2S_2O_3$ = Sodium Hyposulphite.
H_2SO_3 = Hydrogen Sulphite or Sulphurous acid.
H_2SO_4 = “ Sulphate or Sulphuric acid.
$KClO$ = Potass. Hypochlorite.
$HClO_2$ = Hydrogen Chlorite or Chlorous acid.
$HClO_3$ = “ Chlorate or Chloric acid.
$KClO_4$ = Potass. Perchlorate.
Cl_2O = Hypochlorous oxide.

The class of bodies which we have called **acids** are more commonly named in the following manner:—

The negative atom, with the prefixes and suffixes usually attached to the positive element, is named first; this is followed by the word **acid**, thus:—

HCl = Hydrogen Chloride or Hydrochloric acid.
HBr = " Bromide or Hydrobromic "
HI = " Iodide or Hydroiodic "
H_2S = " Sulphide or Hydrosulphuric acid.

The binary compounds, above mentioned, are called **hydracids** to distinguish them from the **oxacids** or ternary acids; and in the naming of these acids the prefix **hydro** is used.

EXAMPLES OF OXACIDS.

$HClO$ = Hydrogen, Hypochlorite or Hypochlorous acid.
HNO_2 = " Nitrite " Nitrous "
$HClO_3$ = " Chlorate " Chloric "
HNO_3 = " Nitrate " Nitric "
H_2SO_3 = " Sulphite " Sulphurous "
H_2SO_4 = " Sulphate " Sulphuric "
H_2CO_3 = " Carbonate " Carbonic "

It will be seen that each of the above acids has a characteristic negative group of atoms; thus in **sulphates** we may always expect the group SO_4''. In a sulphite we will always find the group SO_3''; in a nitrate NO_3'; in a chlorate ClO_3', etc.

Nomenclature Simplified:—The above rules of nomenclature may be applied much more easily by the student, by reference to the table given in Art. 94.

By the use of this table, the student can easily learn to name all the more common compounds. Let it be desired, for example, to name the following formula: $Cu\ SO_4$. Cu is the symbol for **copper**. We next look among the compound radicals and find SO_4 to be the characteristic group of atoms found in all **sulphates**. The name of the formula is, therefore, copper sulphate. What is the formula of zinc carbonate? We find zinc among the dyads, as also the group CO_3, opposite the word carbonates. Both these radicals are dyad, and as a dyad is equivalent to a dyad, they will combine directly, and we shall have $Zn\ CO_3$ as the desired formula.

What is the formula of sodium sulphate? The symbol of sodium is Na; it is monad, while SO_4 is dyad. Two monads are equivalent to one dyad. Hence Na_2SO_4 is the formula of sodium sulphate. Let it be desired to know the formula of calcium phosphate. Here we wish to combine a triad and a dyad. To do so we must double the triad to get an even number of bonds; and then we must take three atoms of calcium to get six bonds. The formula will thus be $Ca_3(PO_4)_2$. What is the formula of Ferric Oxide? In ferric iron two atoms always go together as a hexad. It will require three dyads to saturate the hexad, and we have Fe_2O_3, or for the hydrate $Fe_2(O{-}H)_6$, or Fe_2Cl_6 for the chloride. Stannous chloride will have the formula $SnCl_2$, while stannic chloride will be $SnCl_4$. Mercuric oxide will be HgO, but mercurous oxide will be Hg_2O. Cuprous chloride $= Cu_2Cl_2$.

Examples for Practice.—We introduce here a series of examples for practice in nomenclature, with corresponding names in columns below. The numbers opposite the formulæ will be found opposite the corresponding names below.

1 BaO_2H_2
2 CaC_2O_4
3 $BiCl_3$
4 Na_2CO_3
5 $MgSO_4$
6 $Fe_2(SO_4)_3$
7 $AgNO_3$
8 $(NH_4)Cl$
9 HNO_3
10 Hg_2Cl_2
11 $PbCrO_4$
12 KI
13 K_2O
14 As_2O_3
15 Cu_2FeCy_6
16 $(NH_4)_2S$
17 As_2S_3
18 NH_4MgPO_4
19 $KHCO_3$
20 $SbCl_3$
21 $BiONO_3$
22 $Fe_3(Fe_2Cy_{12})$
23 $K(CN)S$
24 K_2CrO_4
25 $Fe_2(CrO_4)_3$
26 $Al_2(OH)_6$
27 KCy.
28 ZnO
29 $SrCO_3$
30 $Sr(NO_3)_2$
31 $BaCl_2$
32 NH_4NO_3
33 $(NH_4)_2CO_3$
34 $KNaSO_4$
35 $NaHCO_3$
36 $Ca_3(PO_4)_2$
37 $Ca(PO_2H_2)_2$
38 $NaClO$
39 Bi_2O_3
40 O_2
41 $KClO_4$
42 $Pb\text{-}O(C_2H_3O_2)_1H$

1 Barium Hydrate.
2 Calcium Oxalate.
3 Bismuth Chloride.
4 Sodium Carbonate.
5 Magnesium Sulphate.
6 Ferric Sulphate.
7 Silver Nitrate.
8 Ammonium Chloride
9 Hydrogen Nitrate. (Nitric Acid).
10 Mercurous Chloride.
11 Lead Chromate.
12 Potassium Iodide.
13 Potassium Oxide.
14 Arsenious Oxide.
15 Copper Ferrocyanide.
16 Ammonium Sulphide
17 Arsenious Sulphide.
18 Ammon. Magnesium Phosphate.
19 Potassium Hydrogen Carbonate or Potass. Bicarbonate.
20 Antimonous Chloride.
21 Bismuth Oxy-nitrate.
22 Ferrous Ferricyanide.
23 Potassium Sulphocyanate.
24 Potassium Chromate.
25 Ferric Chromate.
26 Aluminic Hydrate.
27 Potassium Cyanide.
28 Zinc Oxide.
29 Strontium Carbonate
30 Strontium Nitrate.
31 Barium Chloride.
32 Ammon. Nitrate.
33 Ammon. Carbonate.
34 Potassium Sod. Sulphate.
35 Sodium Hydrogen Carbonate.
36 Tri-Calcium Phosphate.
37 Calcium Hypophosphite.
38 SodiumHypochlorite
39 Bismuth Oxide.
40 Oxygen.
41 Potassium Perchlorate.
42 Basic Plumbic Acetate.

97. **Irregularities in Nomenclature.**—In many medical and pharmaceutical works, the old style of making the negative precede the positive, with the preposition **of** between them, is used. In this case **per** is used instead of **ic** or **ate**, and **proto** instead of **ite** or **ous**.

EXAMPLES.

	NEW NAME.		OLD NAME.
$SnCl_2$	= Stannous Chloride	or	Protochloride of tin.
$SnCl_4$	= Stannic "	or	Perchloride of tin.
Fe_2Cl_6	= Ferric "	or	Perchloride of iron.
$Fe_2(SO_4)_3$	= " Sulphate	or	Persulphate of iron.
Fe_2O_3	= " Oxide,	or	Per- or Sesquioxide of iron.
Hg_2I_2	= Mercurous Iodide	or	Protiodide of mercury.
Hg_2Cl_2	= " Chloride	or	Protochloride, mild chloride, or calomel.
$HgCl_2$	= Mercuric chloride	or	Bichloride, corrosive sublimate.
HgO	= Mercuric oxide	or	Red oxide of mercury.
Hg_2O	= Mercurous oxide	or	Black oxide of " Protoxide of "

The **proto**-salts of iron are the **ferrous salts**, while the **per** salts are the **ferric** salts.

The names of the oxides of the alkaline metals, the earths, and alkaline earths are sometimes named as follows:—

Al_2O_3	Alumina.	CaO	Lime.
MgO	Magnesia.	K_2O	Potassa or Potash.
BaO	Baryta.	Na_2O	Soda.
SrO	Strontia.		

Some writers name those oxides of the non-metallic elements which dissolve in water to form acids, as though they were formed from the acids by abstracting one or more molecules of water.

Thus SO_2 is named Sulphurous Anhydride.
CO_2 " " Carbonic "
N_2O_3 " " Nitrous "
N_2O_5 " " Nitric "
P_2O_5 " " Phosphoric "
For, $SO_2 + H_2O = H_2SO_3$ = Sulphurous acid.
$CO_2 + H_2O = H_2CO_3$ = Carbonic "
$N_2O_5 + H_2O = 2HNO_3$ = Nitric "
$P_2O_5 + 3H_2O = 2H_3PO_4$ = Phosphoric "
$SO_3 + H_2O = H_2SO_4$ = Sulphuric "

It is a common custom with some authors to use the numerals di, tri or ter, tetra and penta, to indicate the number of atoms of the element to whose name the numeral is prefixed.

Thus:—

	NEW NAME.		OLD NAME.
FeS_2	Ferric Disulphide	or	Bisulphide of Iron.
Fe_2S	Diferrous Sulphide	or	Sulphide of Iron.
Fe_2S_3	Ferric Sulphide	or	Sesquisulphide of Iron.
FeS	Ferrous Sulphide	or	Protosulphide of Iron.
CO_2	Carbon Dioxide	or	Carbonic Acid.
PCl_3	Phosphorus Trioxide,		
PCl_5	" Pentoxide,		
$Hg\ Cl_2$	Mercuric Dichloride,		Bichloride of Mercury.

A few compounds are known by names which do not express their composition.

Thus:—H_3N Ammonia.
CN Cyanogen (symbol Cy).
H_3Sb Antimoniuretted Hydrogen or Stibine.
H_3As Arseniuretted " or Arsine.
H_2S Sulphuretted " or Hydrosulphuric Acid.
H_3P Phosphuretted " or Phosphine.
H_4C Light Carburetted " or Marsh Gas.
H_4C_2 Heavy Carburetted " or Olefiant Gas.

A glossary of obsolete and popular names, and those of some chemical compounds occasionally met with, will be found in the Appendix.

98. **Chemical Reactions and Equations.**—All material bodies, under certain conditions, may undergo marked changes in properties. As the physical properties of bodies depend upon the properties of their molecules, any great change in these properties must depend upon a corresponding change in the molecules. In a homogeneous mass of matter, all molecules are alike; and any chemical change which we are able to produce in one molecule of such a mass, may, with certainty, be produced in all. Hence, by representing the changes which take place between two dissimilar molecules, we may, in reality, represent the changes taking place between the masses of which these molecules form a part. It is upon this principle that we represent chemical changes to the eye. When two substances, on being brought together, act upon each other, the mutual action between them is called a **reaction.**

A body which is added to another to cause such a change, is called a **reagent.** When a jet of coal gas is burned in the air, the reagents are the gas and the oxygen of the air. The **results** of the reaction are light and heat. The **products** of the reaction, are the watery vapor and carbon dioxide which are produced. The **factors** entering into the reaction, are oxygen and the compounds which compose the gas.

While all material molecules are more or less liable to undergo chemical change by the action of external agencies, some do so very readily, while others resist such changes with considerable force. Chemical reactions are favored by anything that lessens cohesion, or favors the free movement of the molecules; as solution, pulverization, trituration, heat, light and electricity.

Reaction between solids is always slow, and, in many cases, entirely wanting. If the solids are brought together in solution, the reaction takes place with readiness; if volatilized, still more readily. Reactions between gases usually take place almost instantaneously throughout the mass, and in many cases, with an explosion. Heat usually favors chemical action, and cold retards it. Light favors many kinds of chemical change, but does not affect all. Reactions, in the laboratory, are generally conducted in solutions. When the bodies are soluble in water, that liquid is generally selected; if not, some other solvent, such as ether, alcohol, chloroform, etc., may be employed.

When two or more substances are brought together in solution, the action that will take place depends largely upon the following conditions, first formulated by Berthelet, and usually known as the **Laws of Berthelet**:—

1st. When two or more substances are brought together in solution, if by any rearrangement of the atoms a product can be formed which is insoluble in the liquid present, that substance will form and separate as a precipitate.

2d. When two substances are brought together in solution, if a gaseous body, or one volatile at the temperature of the experiment, can form, it will form and escape as a gas or vapor.

Illustration.—$BaCl_2 + Na_2SO_4 = ?$ By a rearrangement of these atoms, according to the principles stated in preceding articles, there can only form $BaSO_4$ and $2NaCl$. The latter of these is soluble in water, while the former is not; hence, $BaSO_4$ would always separate from this mixture.

$Zn + 2HCl = ZnCl_2 + H_2$. Here, by changing the places of the two positives, hydrogen is set free, and escapes.

As matter is indestructible, it follows that there can be neither loss nor gain in the weight of the matter taking part in a reaction. The sum of the weights of the factors entering into a reaction must, therefore, be equal to the sum of the weights of the products coming from it. Hence, if we place the sum of the formulæ of the factors equal to the sum of the formulæ of the products of any reaction, it must always form a true equation. In writing out representations of chemical reactions, the student should remember the following rules:—

1st. **Positives** combine with **negatives** and not with **positives.**

2d. Every member of the equation must represent a whole molecule, or a number of molecules.

3d. The quantivalences of the atoms, and radicals must all be saturated according to the rules laid down in article 96.

4th. An **acid** and a **base** cannot exist in the same solution. They are incompatibles, and neutralize each other.

5th. The strongest acid generally selects the strongest base, except in cases where this is modified by Berthelet's laws. Compound radicals, as a rule, remain as such in the products.

To write chemical equations, place the formulæ of the factors, connected by a plus sign, equal to the formulæ of the products, also connected by a plus sign. Now take such a number of molecules as factors that only whole molecules can be produced in the products.

EXAMPLES.

$$\underset{\substack{\text{Silver}\\\text{Nitrate.}}}{\overset{+}{Ag}\,\overset{-}{NO_3}} + \underset{\substack{\text{Sodium}\\\text{Chloride.}}}{\overset{+}{Na}\,\overset{-}{Cl}} = (?)$$

We first determine which are positive and which are negative radicals. The metals are positive (Art. 89), and the non-metallic radicals are negative, as indicated by signs above the symbols.

We next cause the positive radicals to exchange places, whence we have Ag Cl and Na NO$_3$.

On referring to the table of quantivalence (Art. 94), we find all these radicals to be **monad**, and therefore chemical equivalents.

The completed equation will, therefore, be—

$$\underset{\substack{\text{Silver}\\\text{Nitrate.}}}{Ag\,NO_3} + \underset{\substack{\text{Sodium}\\\text{Chloride.}}}{Na\,Cl} = \underset{\substack{\text{Silver}\\\text{Chloride.}}}{Ag\,Cl} + \underset{\substack{\text{Sodium}\\\text{Nitrate.}}}{Na\,NO_3}$$

Complete the following:

1. $2Ag\,(NO_3) + H_2S = ?$
2. $Pb\,(NO_3)_2 + H_2S =$
3. $H_2\,(SO_4) + Ca\,O_2\,H_2 =$
4. $KI + Ag\,NO_3 =$
5. $Fe\,Cl_2 + 2KOH =$
6. $Fe_2\,Cl + 6KOH =$
7. $Ni\,NO + Na_2S =$
8. $Mg\,SO_4 + (NH_4)\,OH =$
9. $Ba\,Cl + Na_2\,SO_4 =$
10. $Bi\,Cl_3 + H_2O = Bi\,OCl + ?$
11. $Pb\,(C\,H_3O_2)_2 + H_2\,SO_4 =$
12. $Ca\,Cl_2 + NH_4)_2\,CrO_4$
13. $Na_2\,CO_3 + 2HCl =$
14. $Cu\,SO_4 + NaOH =$
15. $Hg_2\,(NO_3)_2 + NaCl =$
16. $Mg\,SO_4 + H\,Na_2PO_4 + NH_4OH =$

99. **Stochiometry.**—Chemical symbols represent definite weights, or atomic weights. Chemical formulæ, therefore, enable us to calculate the percentage of any ingredient in the compounds they represent; or, from chemical equations, we may calculate the weight of any substance required by any given process, or the exact amounts evolved by it.

These calculations are all based upon the atomic weights. Molecular weights are derived from the atomic weights.

The molecular weight of calcium carbonate, $Ca\ CO_3$, is

$(C = 12) + (O_3 = 48) + (Ca = 40) = 100$. (See Table, Art. 82.)

On inspection, we see that $\frac{40}{100}$ of the whole quantity are calcium, $\frac{12}{100}$ carbon and $\frac{48}{100}$ oxygen.

Let it be desired to calculate the quantity of hydrogen in one part of water; formula H_2O.

$(H_2 = 2) + (O = 16) = 18 \quad \frac{2}{18} =$ Hydrogen, $\frac{16}{18} =$ Oxygen.

Stated in the form of a proportion, this would be $18 : 2 :: 1 : \frac{2}{18} = \frac{1}{9}$. In this proportion, the fourth term must bear the same relation to the third, that the second does to the first.

Stated in this way, the first problem would be as follows:

$Ca\ CO_3 : Ca :: 1 : x =$
$100 : 40 :: 1 : x = \frac{40}{100} = 40$ or 40 per cent.

In other words, calcium carbonate contains 40 per cent. of calcium. The same calculation may be made for oxygen, as follows:

$Ca\ CO_3 + O_3 :: 1\ x = 48$ or 48 per cent.
$100 + 48 :: 1\ x = 48$ or 48 per cent.

If, instead of one part, we desire the amount in ten parts, we substitute 10 for one in the third term of the equation, thus: $100 : 48 :: 10 : 4.8$ parts.

The fourth term of such an equation must always be of the same denomination as the third.

From the above, we easily deduce the following rule for the statement of such problems: **As the formula of the substance given is to the formula of the substance required, so is the weight of the substance given to x the weight of the substance required.** Reduce the formulæ to their numerical equivalents, and find the value of x.

When three terms of an equation are given, the fourth may be found by multiplying the two **means** (second and third terms), and dividing the product by the given **extreme.**

In calculating the per cent. of any ingredient, by the above rule, the weight given is understood to be 100, *i. e.* per cent. is parts per hundred.

Calculations based upon a reaction may be illustrated as follows:—

Problem.—How much sulphate of zinc can be prepared from 10 grammes of zinc?

Reaction.—$Zn + H_2SO_4 = ZnSO_4 + H_2$

Statement.—$Zn : ZnSO_4 :: 10 : x$

Numerical Statement.—$65 : 65 + 32 + 64 = 161 : 10 : x$

Solution.—$161 \times 10 = 1610 \div 65 = 24.8$ Ans. in grammes.

Problem.—How much $NaNO_3$ will be required to make one pound of HNO_3?

Equation.—$NaNO_3 + H_2SO_4 = NaHSO_4 + HNO_3$

The only terms of this equation concerned in the problem are $NaNO_3$ and HNO_3; the latter being the substance given.

Statement.—$HNO_3 : NaNO_3 :: 1 : x$ pounds.

Numerical Statement.—$63 \div 85 :: 1 : x$

Solution.— $1 \times 85 = 85 \div 63 = 1.35$ pounds.

PART III.

INORGANIC CHEMISTRY.

100. **Classification of the Elements.**—For convenience of study, some system of classification of the elements is necessary. Many systems of classification have been proposed, but all are open to criticism; yet, we may adopt one of these with the understanding that the classification is largely an arbitrary one, and serves merely for convenience. Berzelius was the first to divide the elements into two great classes, to which he gave the names **metals** and **metalloids.** The metals are those elements which possess more or less lustre and opacity, readily conduct heat and electricity, and are electro-positive in combinations.

The non-metals, or metalloids, are such as are gaseous; or, if solid, have no lustre, ductility or malleability, are poor conductors of heat and electricity, and are **electro-negative** in combinations.

This division, while it serves a general purpose, is not capable of exact application; for there are a number of the elements which are **positive** in one combination and **negative** in another. Iodine and arsenic, which most chemists regard as metalloids, have a decided lustre, and the latter forms alloys by fusion with the metals; indeed, there is no line of demarcation, between these two classes, which can be regarded as fixed.

Some classification is necessary, which is not based upon the physical properties alone, but upon their chemical properties; a classification which brings together those elements which have similar chemical properties, and similar compounds with other elements; thus enabling the student to better associate the facts of each in his mind.

There are two important chemical characters upon which most attempts at classification of the elements into groups are based; viz: quantivalence, and electrical polarity of the atoms. By a consideration of both of these properties, the elements may be grouped so as to bring similar elements together.

The behavior of the oxides of the elements with water, may be taken as an index of their polarity. Electro-negatives dissolve in water and form acids, while electro-positives form bases, and some again play a neutral or double rôle.

The most convenient classification of the elements is the one that will bring together into groups those most nearly related in chemical properties. For this purpose the **quantivalence** and **electro-chemical properties** will be employed in this book, with the exception of the two elementary gases, **oxygen** and **hydrogen**. These two elements are taken out of their groups and studied separately, owing to their importance among the elements, and their typical character. The following division of the more important elements seems to be best adapted to the wants of the student:—

THE NON-METALLIC ELEMENTS.

Group I. Hydrogen and Oxygen.
Group II. The **Chlorine Group**, or the **Negative Monads.**
Fluorine, Chlorine, Bromine, Iodine.
Group III. The **Sulphur Group**, or **Negative Dyads.**
Oxygen (by some authors), Sulphur, Selenium, Tellurium.
Group IV. The **Nitrogen Group**, or the **Negative Triads.**
Nitrogen, Phosphorus, Arsenic, Antimony, Bismuth.
Group V. Boron.
Group VI. The **Carbon Group**, or **Negative Tetrads.**
Carbon, Silicon.

THE METALS.

Group I. The **Alkaline Metals.**
Lithium, Sodium, Potassium, Rubidium, Cæsium and Ammonium.
Group II. The Alkaline Earths.
Calcium, Strontium, Barium.
Group III. The Magnesium Group.
Magnesium, Zinc, Cadmium.
Group IV. The Iron Group.
Chromium, Manganese, Iron.
Group V. Nickel Group.
Nickel, Cobalt.
Group VI. Lead.
Group VII. Copper, Mercury.
Group VIII. Silver, Gold.
Group IX. Glucinium, Aluminium, Gallium, Indium.
Group X. The Platinum Metals.
Platinum, Palladium, Rhodium, Ruthenium, Iridium.
Group XI. Molybdenum, Osmium, Tungsten.

TYPICAL ELEMENTS.

HYDROGEN.

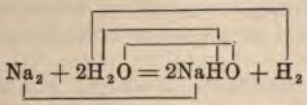
Symbol, H. Atomic weight, 1. Molecular weight, 2. Density, 1. Weight of one litre = 1 crith = .0896 grms. Quantivalence, Monad.

101. **Occurrence.**—Hydrogen occurs in a free state in the gases from volcanoes, fumeroles of Iceland and Tuscany, and in the atmosphere of the sun; in combination, it exists in water, and in most organic substances of both animal and vegetable origin. It is a necessary constituent of all acids and ammoniacal compounds.

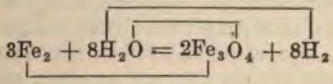
102. **Preparation.**—Hydrogen may be prepared—

First.—By the decomposition of water, by a strong electric current, which splits the water up into hydrogen and oxygen; the former appearing at the negative and the latter at the positive pole.

Second.—By decomposing water by certain metals. When sodium or potassium is used, the decomposition takes place in the cold; but with iron and some other metals at a red heat.

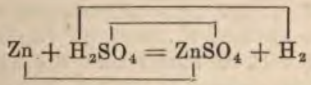


$$Na_2 + 2H_2O = 2NaHO + H_2$$



$$3Fe_2 + 8H_2O = 2Fe_3O_4 + 8H_2$$

Third.—By the decomposition of the mineral acids with some metal, as zinc, iron, or magnesium. In this case, the metal takes the place of the hydrogen, which is crowded out of the molecule of the acid.



$$Zn + H_2SO_4 = ZnSO_4 + H_2$$

Water is added to dissolve the zinc sulphate formed, and to prevent it from crystallizing on the surface of the zinc. Chemically pure zinc, however, will not dissolve in very dilute acid, unless it be made one pole of a galvanic couple.

This method is the one usually employed for the preparation of hydrogen in the laboratory. The gas prepared from commercial zinc and acid is not pure, however, as it contains other gases derived from impurities in the materials used. Pure hydrogen, in small quantities, may be prepared by the first method, or by decomposing water with an alloy of sodium and mercury—sodium amalgam.

103. **Properties.**—When pure, at ordinary temperatures and pressures, hydrogen is a colorless, transparent, odorless, tasteless gas. It is 14½ times lighter than air, being the lightest gas known.

One litre of it at 0°C. and 760 mm. pressure, weighs .0896 grams. = the crith. It is almost insoluble in alcohol, and at a temperature of —140°C and under a pressure of 650 atmospheres, it has been condensed to a steel-blue liquid. It is the best conductor of electricity and heat among the gases. It is very diffusible, and a vessel to contain it, must be made of glass or some very compact material. Certain metals absorb large quantities of it. Palladium will absorb 900 times its volume of the gas: spongy platinum, sodium, potassium and iron also absorb considerable quantities of it. This action of the metals is called **occlusion.** During the condensation of the gas in the pores of the metal, the latter is heated to a considerable degree. A jet of hydrogen may be ignited by directing it upon a ball of finely divided platinum, or a ball of asbestos, which has been dipped into a solution of platinum chloride, and heated in the flame of a lamp.

Under ordinary conditions, hydrogen has little tendency to unite with the other elements, chlorine being the only one with which it combines directly, and then only under the influence of light. At higher temperatures it unites with oxygen, and is, therefore, readily combustible in the air, burning with a bluish and very hot flame. The product of the combustion is watery vapor. A given weight of hydrogen produces more heat in burning than any other known combustible. It will not maintain animal respiration, but is not poisonous. A lighted taper is extinguished on being thrust into it. If hydrogen and oxygen be mixed together and a lighted taper applied to them, an explosion takes place. On examination, it will be found that the hydrogen has combined with one-half its volume of oxygen, and if there be present more than this quantity of the latter gas, it will remain undisturbed. If, on the other hand, there be less than one-half as much oxygen as hydrogen, there will remain a part of the hydrogen. In other words, when these gases combine they do so only in the one proportion. This law holds good with all chemical combinations, and is known as the **law of definite propor-**

tions. Hydrogen has so much of a tendency to unite with oxygen at high temperatures, that it will take it from many metallic oxides, and leave the metal in the free state.

This process is called **reduction** or **deoxidation.** It is by this process that the **reduced iron** or **iron by hydrogen** of pharmacy is produced.

Hydrogen will unite quite readily with some elements which it ordinarily does not affect, if they be put into the flask where the hydrogen is generating. Arsenic and antimony compounds, for example, are split up, and these elements unite with the hydrogen. Many oxides are reduced, and chlorine is removed from chlorides under these circumstances. The greater energy of the hydrogen, in these cases, is explained by the supposition that at the moment of liberation of the hydrogen atoms, and before they have combined into molecules, they are ready to take up with any atom with which they may come in contact, and are stronger in their affinities, before combining with a neighboring hydrogen atom, by just the force it will take to decompose the hydrogen molecule when once formed. This condition of an element is known as the **nascent state** (from **nascere,** to be born). When any chemical action takes place between molecules, there is a considerable expenditure of force required to break up the combinations already formed, before new ones can be formed; and when these combinations do not exist, the new combinations take place with ease. Hydrogen is one of the constituents of the gases of the stomach and intestines, and is frequently found in the gases exhaled from the lungs. Its physiological properties, if any, are slight.

In its chemical properties, hydrogen resembles the metals more than the metalloids, usually playing the positive rôle, and forming salts in which it occupies the place of metallic atoms in similar compounds, and is easily substituted for them, or displaced by them. On this ground, the acids may be regarded as salts of hydrogen.

104. **Uses.**—The uses of hydrogen are limited. Owing to its lightness, it is sometimes used to fill balloons. The ascensional power, or the lifting power, of one litre of hydrogen, is found by deducting its weight, .0896 grms., from the weight of one litre of air, 1.2932 grammes, or 1.2036. The lifting power of one cubic foot is about 525 grains, or 1 ounce and 55 grains. Hydrogen is also used with oxygen in the oxy-hydrogen blowpipe. In the laboratory, it is used as a reducing agent.

OXYGEN.

Symbol, O. Atomic weight, 16. Molecular weight, 32. Density, 16. Weight one litre, 1.43 grammes.

105. **Occurrence.**—Oxygen was discovered by Priestley, in 1774, in the air, of which it constitutes 20.6 per cent. It exists in the air in the free or uncombined state, mixed with nitrogen and small quantities of other gases. It enters into the composition of a great variety of compound bodies, such as minerals, vegetable and animal bodies. Water is eight-ninths, sand one-half, and alumina one-third oxygen, by weight.

106. **Preparation.**—Oxygen may be prepared:—

First.—By heating mercuric oxide in a retort or flask, when it breaks up into oxygen and black mercurous oxide; or, if the temperature be high, into oxygen and metallic mercury.

$$2HgO = 2Hg + O_2.$$

Second.—By heating black manganic oxide (MnO_2) to redness, in an iron or clay retort, when it gives off a part of its oxygen.

$$3MnO_2 = Mn_3O_4 + O_2.$$

Third.—By decomposing acidulated water with a current of electricity. The oxygen obtained in this way is very pure, but it is too slow for ordinary use.

Fourth.—The best method, and the one generally employed, is by heating potassium chlorate.

$$2KClO_3 = 2KCl + 3O_2.$$

The evolution of the gas takes place more regularly and at a lower temperature, if the chlorate be mixed with ferric oxide, cupric oxide, or manganic dioxide. In practice, the last is generally used in the proportion of one part of the oxide to two or three parts by weight of the chlorate. The manner in which the oxide acts is somewhat obscure, for it seems to undergo no change in composition, and is found to be unaltered in the residue left in the retort.

The process may be conducted in a round-bottom glass flask, furnished with a large-size delivery tube, provided that the heat is carefully regulated and not allowed to become too high.

One kilogramme of the chlorate ought to yield about 140 litres of oxygen.

107. **Properties.**—Oxygen, when pure, is a colorless, transparent, odorless, tasteless gas, slightly heavier than air. Its sp. gr. is 1.10563. Water dissolves three per cent. of its volume,

at ordinary temperatures. Under a pressure of 300 atmospheres, at a temperature of —140°C, it condenses to a transparent liquid whose specific gravity is 0.9787 (Pictet). It is magnetic. The magnetism of the atmospheric oxygen is equal to that of a layer of iron covering the surface of the earth 0.1 millimetre in thickness.

Oxygen forms oxides of all the known elements except fluorine. Its range of affinities, and its energy of combining power are its characteristic chemical properties. Most elements combine directly with it, especially at high temperatures. When this oxidation is accompanied by light and heat it is called **combustion.** A body is said to be combustible when it unites readily with oxygen, and liberates light and heat in so doing. A combustible body usually requires to be heated to a more or less elevated temperature before it will be acted upon by the oxygen; but when the process has once begun it is kept up by the heat generated in burning.

Some bodies, not usually regarded as combustible, will burn when heated to a red heat and plunged into an atmosphere of pure oxygen; as, for example, a steel watch spring or small iron wire, so treated, will burn with great brilliancy. Bodies which burn in air with difficulty, burn in pure oxygen with great readiness. Oxygen is the great supporter of combustion, but the action of oxygen and the combustible body are mutual. A jet of air may be burned in a jar of illuminating gas or hydrogen as readily as these last burn in the air. Oxidation often takes place slowly, and the heat produced, although the same in both cases, passes off into the air or surrounding bodies, so that the temperature does not rise much above that of the air. This is sometimes termed **slow combustion**; or, more commonly, oxidation.

Most ordinary combustibles contain carbon and hydrogen, and in such cases the results of the process are carbonic dioxide and watery vapor. In case the combustible contains sulphur, it becomes sulphurous oxide; if nitrogen, it becomes either ammonia or free nitrogen, according as the oxidation is complete or incomplete. The respiration of animals is very similar to combustion.

108. **Uses.**—The uses of oxygen are very many. The oxygen taken into the air vesicles of the lungs passes through their thin walls, by diffusion, into the blood. There it combines with the hæmoglobin and circulates with it throughout the body, assisting in burning up the waste products of the broken down tissues.

One hundred volumes of arterial blood from a dog contain twenty-two volumes of oxygen (Grehaut), and this quantity varies with the amount of hæmoglobin, or with the red globules.

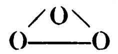
In the air, the oxygen is mixed with nitrogen to dilute it and regulate its action. Oxygen is of use as a supporter of combustion, to afford us artificial heat and light. With this heat we drive our steam engines, warm our houses, smelt our ores and cook our food.

109. **Ozone.**—If a series of electric sparks be passed, for a few minutes, through a portion of air or oxygen gas confined in a tube, it acquires a peculiar pungent odor, and exhibits properties which it did not previously possess. The same odor is usually detected in the air in the neighborhood of a frictional electrical machine while in operation; or in the gas given off by a mixture of potassium permanganate and sulphuric acid; or when phosphorus, partially covered with water, is exposed to the air and allowed to undergo slow oxidation; or by the electrolysis of water containing sulphuric and chromic acids. Ozone can often be detected about a galvanic battery, using as the exciting fluid a solution of sulphuric acid and potassium bichromate.

Ozone has not been prepared in the pure state. The highest degree of concentration reached does not exceed one per cent.

110. **Properties.**—The properties of ozone are those of oxygen, intensified. It is a very powerful oxidizing agent, tarnishes silver and mercury, sets iodine free from potassium iodide, and is readily destroyed by contact with easily oxidizable organic matters, and by a temperature of 260°C. In this last case, it is reconverted into oxygen. It is a strong bleaching agent. It is soluble in oil of turpentine and in ether.

In preparing ozone from oxygen, a contraction takes place, and it again expands on being reconverted into ordinary oxygen. This shows that it is a condensed form of oxygen.

Ozone is 1½ times heavier than oxygen, and its molecule is represented by  or O_3.

111. **Tests.**—The presence of ozone may be detected by its action upon a paper saturated with a solution of potassium iodide and boiled starch paste. This paper becomes blue by its action, owing to the liberation of iodine, which gives a blue color with starch. A piece of reddened litmus paper saturated with potass. iodide is also blued. A paper moistened with an alcoholic solution of guaiacum is also changed to a light blue by its

action. A piece of paper impregnated with a solution of manganous sulphate or lead hydrate turns dark brown or black in its presence. These reactions disappear when the air is heated to 260°C. (500°F.). Ozone is found in the air, especially after thunder storms, and when in appreciable quantities acts as a purifier of the air, destroying, by its oxidizing action, many forms of organized germs hurtful to animal and vegetable life. On this account, it has been regarded as a valuable antiseptic and disinfectant. As it is very irritating to the mucous membranes, and when present to any considerable extent causes distressing coryza or even hæmoptysis, it is to be recommended with caution.

HYDROGEN AND OXYGEN.

Two compounds of these elements are known.
Hydrogen oxide, or water, H_2O.
Hydrogen peroxide, or oxygenated water, H_2O_2.

HYDROGEN OXIDE, OR WATER.

112. **Occurrence.**—Water is so widely distributed in nature, that it is almost universal. It exists in the three states of solid, liquid, and gas or vapor.

It occurs in the solid form, below the temperature of 0°C. (32°F.), and as liquid between 0°C. and 100°C. (212°F.). In the form of vapor it exists in the air at ordinary temperatures. It is poured into it from combustion, in various manufacturing processes, from volcanoes, by spontaneous evaporation from the surface of the ground, bodies of water, and leaves of foliage. Seven-eighths of the entire human body is water. Potatoes contain 75 per cent.; water-melons, 94 per cent. and cucumbers 97 per cent. of water. It enters intó the composition of many rocks, and forms a necessary part of many crystals, where it is known as water of crystallization.

113. **Composition.**—The composition of water may be determined in two ways: by analysis or by synthesis. If a current of electricity be conducted through a vessel of water, slightly acidulated with sulphuric acid, the water will be decomposed into the two gases, hydrogen and oxygen, in the proportion of two volumes of the first to one of the second. If, now, these gases be mixed together in the same proportion, and an electric spark sent through a portion of the mixture, they recombine with an explosion. If **equal volumes** of the two gases be

used, there will remain, after the explosion, one-fourth as much gas as was taken, which, on testing, will be found to be oxygen.

These two experiments show that the proportion of the gases, by volume, must be two of hydrogen to one of oxygen. On weighing the two gases, we find that the oxygen weighs eight times as much as the hydrogen; or, by weight, water is composed of $\frac{8}{9}$ oxygen and $\frac{1}{9}$ hydrogen.

The recombination is conducted in the endiometer of the construction represented by Fig. 31 and the decomposition in the apparatus shown in Fig. 30. For the synthesis of the water, the mixed gases are inserted into the graduated limb through the stop-cock at the top. The wires connected with a small induction coil, are connected with the two platinum wires soldered into the glass just below this stop-cock, and which are separated within the tube by a space about $\frac{1}{16}$ of an inch. On connecting the coil with the wires from the battery a spark is sent across the space between the platinum wires, which ignites the gases. In Fig. 30 the wires are seen to pass through the tubes and terminate in two

FIG. 30.

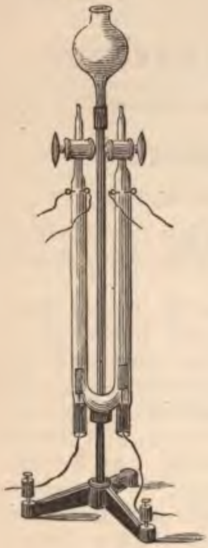
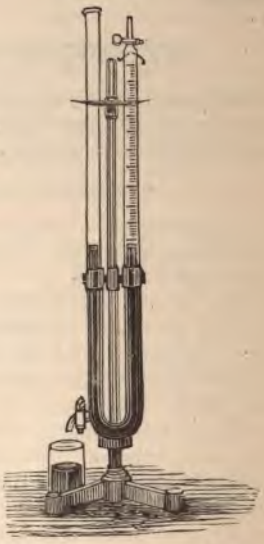


FIG. 31.



strips of platinum foil, from which the gases escape to the top of their respective limbs. If the graduated limb in Fig. 31, containing the gases to be combined, be surrounded by a larger tube through which steam from a kettle is kept passing, and the gases are measured before and after the explosion, at the same temperature, it will be found that the steam produced by the combination of the oxygen and hydrogen, will occupy two-thirds of the volume of the mixed gases before the explosion. That is, the two volumes of hydrogen and the one volume of oxygen have formed two volumes of steam or vapor of water. Applying the law of Avogadro, it will be seen that there are the same number of molecules of water produced as there were of hydrogen taken. Now, as we have seen in Part II of this book, the hydrogen molecule contains two atoms, and hence the molecule of water must contain two atoms of hydrogen. By the same reasoning, it may be shown that the molecule contains but one atom of oxygen; or, the formula is H_2O.

114. **Preparation.**—Water may be prepared in several ways, by chemical means.

First.—The simplest method is the direct union of the gases.

$$2H_2 + O_2 = 2H_2O.$$

Second.—It is always produced by burning hydrogen gas or any combustible containing it, in the air, and may be condensed by conducting the products of the combustion through a tube or flue kept cool by immersing it in cold water.

$$\underset{\text{Marsh Gas.}}{CH_4} + \underset{\text{Oxygen.}}{2O_2} = \underset{\substack{\text{Carbon}\\\text{dioxide.}}}{CO_2} + \underset{\text{Water.}}{H_2O}.$$

Third.—Water is produced as one of the products of the action of an acid upon a base, or metallic oxide. Thus:—

$$\underset{\text{Potass. hydrate.}}{K—O—H} + \underset{\substack{\text{Hydric}\\\text{nitrate.}}}{HNO_3} = \underset{\substack{\text{Potass.}\\\text{nitrate.}}}{KNO_3} + \underset{\text{Water.}}{H_2O}.$$

Fourth.—The reduction of a metallic oxide by hydrogen or some organic substance containing it.

$$C_2H_6O + 6CuO \text{ at red heat} = 3Cu_2 + 2CO_2 + 3H_2O$$
$$2CuO + 2H_2 = Cu_2 + 2H_2O.$$

115. **Properties—Physical.**—When pure, water is a colorless, transparent, mobile liquid, without taste or odor. When viewed in large quantity, however, it has a bluish color. It is a poor conductor of heat and electricity. When water is cooled down below 0°C. (32°F.) it assumes the solid state, called ice.

When the temperature is raised to 100°C. (212°F.) in ordinary conditions, it assumes the gaseous state, called steam. This point is called the boiling point. (See chapter on Heat.) The boiling point is higher than 100°C. under an increased pressure, or when it contains considerable solid matter in solution. Water, at a temperature of 4°C. (39.2°F.), is taken as the unit of specific gravity of liquids and solids. At this temperature it possesses its greatest density. When it is heated above or cooled below this point, it expands and becomes less dense. Water is 773 times heavier than air at 0°C., and 11,147 times heavier than hydrogen. Water expands quite rapidly and with great force, on solidifying, and hence ice is lighter than water. This expansion is supposed to be due to the greater space required for the molecules in arranging themselves into crystals. The form of the crystal of water, is hexagonal. It may frequently be seen in small snow flakes received upon a dark surface; the lines of the three equal axes can often be seen with great distinctness.

The variations in the boiling point of water are much greater than those of the freezing point, but this last is subject to slight variations of temperature. Water may be cooled in capillary tubes to —15° C. (5°F) before it solidifies, if the tubes remain at rest; but if they are agitated, when at this low temperature, the water will instantly solidify. The agitation favors the movement of the molecules into the position to form crystals, and hence large bodies of water freeze at a higher temperature when agitated by a gentle breeze than when the air is very calm. Although water is converted into vapor most rapidly at 100°C., water (even ice and snow) undergoes evaporation at all temperatures, especially when the air is dry. Owing to its great solvent power for solids, pure water is never met with in nature. There are comparatively few substances which are totally insoluble in water. When we wish to prepare pure water, we generally resort to the process of distillation, rejecting the first 20 per cent. of that which distills over, and also the last 20 per cent. It is by no means an easy matter to prepare absolutely pure water, even by this process; but, by conducting the process carefully, with the above precautions, we may obtain a water pure enough for all ordinary chemical purposes. Pure water is generally selected as the solvent of chemical substances which are to be submitted to any chemical change, because the reactions take place more readily in solution than when in the solid state, and because water is a neutral body, which does not complicate the result by taking part in the action, itself. The vapor of water is transparent, invisible and colorless. Its density is 9, and its sp. gr., referred to air, is 0.6234. One volume of water will produce 1696 volumes of steam, or, approximately, one cubic inch of water will produce one cubic foot of steam at 100°C., (212°F.) and absorbs 536.0 units of heat.

Chemical Properties.—We have already referred, in speaking of the preparation of water, to some of its chemical properties. It unites directly with many metallic oxides to form bases or hydrates, and to some oxides of the metalloids to form acids.

$$CaO + H_2O = CaO_2H_2$$
$$SO_3 + H_2O = H_2SO_4$$
$$CO_2 + H_2O = H_2CO_3$$

It enters into a feeble union with many metallic salts in solution, and separates with them when they crystallize, as water of crystallization. Certain substances exhibit a marked tendency to combine with water, or to absorb it from the air, and are used in the laboratory as drying agents. Among these are **calcium chloride, sulphuric acid** and **phosphoric pentoxide.**

116. **Natural Waters.**—As already stated, natural waters are never free from dissolved impurities. They contain gaseous, liquid and solid impurities, varying according to the source from whence derived, the temperature, the nature of the soil or rocks over which it has flowed, or the state of the air at the time. Natural waters may be divided into **potable** (or drinkable), **mineral** and **saline.**

117. **Potable Waters.**—To this class belong well and spring waters, river water, lake water and ice water.

The purest natural waters, are **rain** and **snow water** from mountainous and country districts. The purity of rain water varies with the locality where it falls. In the neighborhood of large cities, where the air is charged with the products of large factories, etc., it will contain whatever of these can be washed out of the air.

Sulphuric acid, for example, is comparatively abundant in the air of large cities. The rain water of London, as given by Dr. R. Angus Smith, contains 20.5 parts per million, while that of inland places in England contains only 5.5 parts; and that from inland places in Scotland, only 2. parts; while from Glasgow it contained 70. parts.

The source of the sulphuric acid is mostly from the combustion of coal containing sulphur. The chlorides in rain water, principally **sodium chloride,** vary with the distance from the sea coast.

Ammonium nitrate and nitrite are found in small quantities, derived from decomposing organic matter, and from the combustion of coal. Another source of these compounds, is the oxidation of a small quantity of the nitrogen of the air by ozone generated by lightning in its passage through it. Rain water also contains more or less dust and organic matter, which it washes out of the air in falling. The gases found in rain water are carbon dioxide (CO_2), nitrogen, oxygen, and sometimes, in cities, sulphur dioxide (SO_2) and hydrogen sulphide (H_2S).

Peligot's analyses show $\left.\begin{array}{ll} CO_2 & 2.4 \\ \text{Oxygen} & 6.59 \\ \text{Nitrogen} & 14. \end{array}\right\}$ c.c. per litre.

It will be noticed that the proportion of oxygen in the air of rain water, is about twice as great as that of the atmosphere.

Rain water, as ordinarily collected on roofs of houses, is very much contaminated with both organic and mineral matter washed from the roof on which it falls. It is very liable to become putrid from the decomposition of this organic matter, and to breed the larvæ of certain insects. Melted snow furnishes a water even purer than rain water, especially if we collect that which falls toward the end of a storm.

Ice water varies very much in purity, according to the purity of the water from which the ice is obtained. It is always purer than the water from which it is formed, and when obtained from clear lakes or rivers, it is often the purest of natural waters, owing to the fact that in crystallization of water, or freezing, it leaves the dissolved solids and gases almost entirely in the remaining water. The absence of the usual gases, however, renders ice-water **flat** to the taste.

Spring and Well waters are simply rain water which has been filtered through a more or less thick layer of soil. The nature and quantity of the dissolved matters will depend upon the nature of the soil and rocks through which it percolates, or over which it flows.

In large cities, where the soil is saturated with filth, the well waters are very impure; while in well drained and mountainous country districts they are much purer. Dangerous organic matter may filter through many feet of soil and poison the water of a well or spring. Shallow wells usually contain much more organic and less mineral matter than deep wells, and are therefore more likely to contain dangerous or unwholesome matters. Wells are essentially a pit for the reception and accumulation of the drainage from the surrounding soil. For convenience they are usually situated near the dwelling, where the soil receives more or less household waste of various kinds, and are often placed near a cesspool or privy vault. The effect of the geological character of the soil is almost entirely obliterated by this local impurity. Such waters, even when disgustingly impure, are usually bright, sparkling and palatable. The saline matters held in solution are often much approved by those accustomed to their use. Deep wells may be regarded as those which draw their supply from a depth of 100 feet or more from the surface. In cases where the supply is drawn from below a dense bed of clay or of impervious rock, the well may be considered deep, when such supply is much less than 100 feet below the surface. Deep wells may be regarded as artificial springs, as both are subjected to the same conditions.

Artesian wells are artificial springs formed by boring into the earth in a low-lying district surrounded by high ground, until a layer of rock or gravel is reached, situated between two impermeable layers and containing water. The strata must be so curved that their outcrop is on a higher plane than the surface of the well. In such cases the water rises to the surface without pumping, and its character is determined by the nature of the rocks in which the water is found.

Surface Waters.—These comprise river water, pond, lake and sea water. The water supply of large cities is usually taken from this class of waters, and consists of spring water, and rain water which has fallen upon a considerable surface of country. Surface waters usually contain a large proportion of organic and mineral matters. Surface water, draining from a cultivated district, contains more organic and mineral matter than that from uncultivated districts, and the character of it is considerably influenced by the application of fertilizers to the land. River waters are often contaminated by the practice of discharging into them sewage and refuse from various manufactories along their banks.

Characters of a Good Drinking Water.—1st. It should be clear and limpid. Cloudy and muddy waters should be avoided. 2d. It should be colorless. A greenish or yellowish color is usually due to vegetable or animal organisms. 3d. It should be odorless; especially free from sulphuretted hydrogen or putrefactive animal matter. 4th. It should not be too cold, but should have a temperature of from 8°C.(46 F.) to 15°C.(60 F.). 5th. It should have an agreeable taste; neither flat, salty, nor sweetish. A certain amount of hardness and dissolved gases give a sparkling taste. It should contain from 25 to 50 c.c. of gases per litre, of which 8 to 10 per cent. is carbonic acid, and the rest oxygen and nitrogen. The air of natural waters is richer in oxygen than the atmosphere above them, viz.: about 33 per cent. of oxygen and 67 per cent. of nitrogen, when the water is fully saturated, which is not always the case. Highly contaminated waters usually contain less oxygen than the above proportion, because it is used up in oxidizing the organic matter. 6th. It should be as free as possible from dissolved organic matter, especially of animal origin. 7th. It should not contain too great an amount of hardness. A certain quantity of saline matter is necessary, however, to give it a good taste. It should not contain over 3 or 4 parts of chlorine in 100,000 parts of water.

118. **Water Analysis.**—We have seen that natural waters are never pure, but contaminated by various kinds of foreign matter to which they have been exposed. These impurities may be harmless to the human economy, or they may be very harmful. It is the object of the analyst to determine, as nearly as possible, the nature and amount of the impurities found in a water, so as to form an opinion as to its healthfulness.

It has been proven, beyond doubt, that water is a fruitful disseminator of disease. It is believed that the disease produced in such cases is due to microscopic organisms, and that the organic matter of the water simply furnishes a suitable solution in which they may live and grow. It is also known that these organisms grow best in solutions of animal organic matter; hence, waters contaminated with animal matters are looked upon as much more dangerous than those containing vegetable matter.

If the above idea be correct, it is evident that chemical analysis cannot detect the disease-producing element, although it can tell **pure** from **impure** water.

Several methods are in use among chemists for the purpose of forming an opinion as to the character of drinking waters. The elements usually relied upon in a sanitary examination, are total residue left on evaporation, the loss in weight of this residue on ignition at a dull red heat, the chlorides, the nitrates and nitrites, ammonia, organic carbon and nitrogen, and the quantity of oxygen the water will absorb from an acid solution of potassic permanganate. Poisonous metals should, of course, be looked for where there can be any suspicions of their presence.

119. **Total Residue.**—The amount of residue left on evaporation serves, in a crude way, to indicate the amount of solids, but is only approximate and of no significance unless it reaches more than 30 or 40 grains per gallon. This residue will, of course, vary with the hardness, or the amount of soluble constituents in the soil. By igniting the solids (heating them over a lamp) we may expel all volatile and organic matter, including the water held by the calcium sulphate, at the temperature of 100°C., (the temperature generally used in drying the residue), some carbon dioxide from carbonates, and oxides of nitrogen from nitrates and nitrites. The residue left after ignition gives a rough idea of the mineral matters present, and the loss should never reach 50 per cent. of the total residue.

120. **Chlorine** in potable waters is very largely derived from the sodium and potassium chlorides of urine and sewage. The average amount of chlorine in urine is not far from 5 parts per 1000, or 500 parts in 100,000 parts of water. The average found in sewage is about 11.5 parts per 100,000. Over 5 parts of chlorine per 100,000 may be considered, in most cases, to be due to pollution of the water by sewage or animal excretions. Proximity to the sea does not materially affect the amount of chlorine, in well waters. Too much dependence cannot be put upon the amount of chlorine in a water, as a means of judging of its purity, for vegetable matter may exist in dangerous quantity without its presence being indicated by the chlorine present. Chlorine is estimated by a standard solution of argentic nitrate.

121. **Hardness.**—The hardness of water is produced principally by the acid carbonates, and the sulphates of calcium and magnesium. The acid carbonates are decomposed by boiling, and the neutral salts precipitated:—

$$CaH_2(CO_3)_2 = CaCO_3 + H_2O + CO_2.$$

Hardness due to these salts is called "temporary," while that due to the sulphates of these metals is called "permanent."

The hardness of a water seems to have little or no influence upon the health of those who use it. Hardness is estimated by means of a standard solution of soap. Temporarily hard waters may be softened by adding lime in sufficient quantity to neutralize the excess of carbonic acid in the water, when the carbonate ($CaCo_3$) or ($MgCo_3$) is formed, which settles to the bottom. The alkaline carbonates (Na_2Co_3 and K_2CO_3) may be used to precipitate permanent hardness and soften the water. Wood ashes, which contain the latter of these salts, is frequently used by washerwomen to soften water for washing.

122. **Nitrates and Nitrites.**—These salts are usually looked upon as evidences of former contamination of a water by nitrogenous organic matter. The decomposition of organic matter of animal origin, in waters containing dissolved oxygen, yields nitrous and nitric acid, which combine with bases present to form salts of these acids. Rain water contains a small quantity of these acids, but a larger quantity indicates that the water is undergoing, or has undergone, a natural process of purification from animal matters. The quantity found usually indicates the amount of matter thus decomposed. This purifying process goes on more slowly in river and lake waters than in well waters, because the exposure to oxygen in the latter is more complete while filtering through the soil, than in a body of water. Deep wells may be allowed to contain more nitrates than shallow ones, for the organic matter may all be destroyed by the filtering process through the soil into a deep well, while in the shallow one some organic matter, germs, and spores capable of causing disease, may pass undecomposed through the shorter distance. A careful estimation of nitrates gives, therefore, considerable knowledge of the past history of a water, and is regarded by chemists as of great importance.

A very easy and delicate test for nitrates and nitrites in water, is the following: Make a solution of *diphenylamine* in pure, strong sulphuric acid, adding a little pure water to make a clear solution.

To the suspected water add some of the sulphuric acid, then a few drops of the above solution. If nitrates or nitrites are present, a deep blue solution will be formed at once. This test may be used by any one, and will give some idea of the safety of a drinking water.

123. **Moist Combustion.**—By this is meant the oxidation of the organic matter found in a water by adding to it a measured quantity of potassium permanganate, with some sulphuric acid, and determining the oxygen absorbed from this salt by the organic matter present. This method cannot distinguish between vegetable and animal matters, neither will it measure the oxygen consumed in oxidizing nitrites to nitrates, or ferrous to ferric salts. In the presence of these salts it may give erroneous results. Most chemists regard the process as of some value in giving an idea of probable pollution by organic matter, as it has been shown that dangerous organic matter is more easily oxidized than that which is harmless.

124. **Argentic Nitrate Test.**—It has been proposed to estimate the amount of organic matter in water by the addition of argentic nitrate, and, after filtering out the precipitated chloride and sulphide of silver, to set the filtered water in a strong light for twenty-four hours. The organic matter present in the water, reduces a portion of the silver to the metallic state, and forms a black deposit. This may be filtered out and weighed, and from the quantity, we may judge of the probable organic purity of the water. As a qualitative test, this may easily be performed and give proximate results.

125. **Ammonia.**—The spontaneous decomposition of organic matter in water first affords ammonia, then nitrites, and finally nitrates. This fact is so generally conceded as to make an estimation of the ammonia found in a water, a very important part of the sanitary examination.

By distilling a measured portion of the water under examination, all the ammonia is found to pass off and condense with the first portions distilled. To determine the amount of ammonia present in the distillate, it is only necessary to compare the color produced in it by a small quantity of Nessler's reagent,* with the color produced by a known quantity of ammonia added to some pure water in another tube.†

After having freed the water from ammonia, as above, if we add to it a strongly alkaline solution of potassium permanganate and continue the distillation, the nitrogenous organic matter present gives up its nitrogen in the form of ammonia, which is denominated "albumenoid ammonia." This may be determined in the same manner as the "free" ammonia above. This method of examining a water for nitrogenous matters, is generally looked upon as one of the best methods in use for the purpose.

126. **Organic Carbon and Nitrogen.**—Drs. Frankland‡ and Armstrong, in 1868, proposed to burn the organic matter in the residue left on evaporating the water to dryness, and by collecting and weighing the nitrogen and CO_2 produced, to estimate the amount of nitrogenous matter present. This process is difficult to conduct, and is open to serious errors; it is, therefore, little employed.

127. **Living Organisms.**—It has lately been proposed to

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* Nessler's reagent is a strong solution of potass. iodide, saturated with mercuric iodide, and rendered alkaline with sodic hydrate.

† See Water Analysis by Wanklyn. London, 1876, p. 32.

‡ Water Analysis by Frankland. Philadelphia, 1880.

test "organic vitality" of waters designed for drinking purposes, by adding enough gelatin solution to make the water slightly viscid, place it in a tall cylinder and allow it to stand in a light place from twelve to twenty-four hours. The gelatin prevents the free movement of the organisms in the water, and allows us to see their growth, by the opalescent spheres caused by great numbers of them congregated in a small spot, and springing from a single parent. By the number and rapidity of the growth of these spheres, we may form some idea of the organic vitality or the power of the water to grow organisms. This method may be modified by floating upon the water, contained in a tall narrow cylinder or test tube, a thick solution of gelatin, and then covering the sides of the tube with paper, admitting light by a small hole in the paper opposite the layer of gelatin.

On standing from 12 to 24 hours in a light place, the organisms existing in the water will be found entangled in the gelatin film, and may be removed with a pipette and placed under the microscope. This process aims to judge of the character of the water by the character of the organisms found growing in it. If large numbers of putrefactive bacteria be found, it would be evidence of putrefying organic matter. We have thus briefly outlined the methods in use for testing drinking waters, for the purpose of giving the student some knowledge of the subject, without going into tedious details, which are intended only for the chemist, and which do not come within the scope of this work.

128. **Purification of Water.**—Water may be separated from suspended impurities by filtration, *i. e.*, by passing it through any porous substance not soluble in the water, as clay, sand, charcoal, brick, unglazed earthenware, unsized paper, etc. For filtering large quantities of water for cities or manufactories, sand or brick is to be preferred. For filtering water for family use, a brick partition two or four inches thick, built in a cistern, works well. The water is delivered in one apartment and pumped from the other for use. Or, a barrel in the bottom of which several holes are bored, may be filled with alternate layers of gravel, sand and charcoal, and placed over the mouth of the cistern or reservoir, through which the water may be filtered. Such a filter, when freshly made, will remove a part of the dissolved organic matters as well as suspended matters. The **silicated carbon** filters found in the market, made of pulverized carbon and cement, are much neater, and will also oxidize a portion of the dissolved organic matter in passing the water through

them, especially when new. Porous stone filters, made from a silicious stone, are to be found in the market, which filter rapidly and very perfectly. Spongy iron filters, made by roasting hematite with coal, are still more active in destroying organic matter and putrefactive germs. We have already referred to methods of precipitating the **hardness** from water. Distillation, as a means of purifying water, has also been referred to. Freezing purifies water, and removes, to a considerable degree, the mineral as well as organic matters; but freezing cannot make a dangerous water perfectly safe. Numerous instances show that disease may be communicated as surely by ice, as by unfrozen water.

129. **Mineral Waters.**—Under this name are included such waters as, from some dissolved substances, have a greater or less therapeutical value. These vary so much in the character of the dissolved substances, that no exact classification of them can be made; but they may be roughly classified as follows .—

1st. **Carbonated waters,** those which are charged with carbonic dioxide or carbonic acid.

2d. **Sulphuretted Waters,** those which contain sulphides of hydrogen or one of the alkaline metals, in notable quantities. They are used for baths as well as for drinking.

3d. **Alkaline Waters,** those containing considerable quantities of carbonates or bicarbonates of the alkaline metals—sodium, potassium, or lithium.

4th. **Saline Waters,** those containing the neutral salts, such as the chlorides, bromides, or iodides of the alkalies or alkaline earths.

5th. **Chalybeate Waters,** those containing some one of the compounds of iron. Closely allied with these in properties, are those containing manganese.

6th. **Thermal Waters,** or such as come to the surface at a temperature above that of 20°C. (68° F.). Some of these springs contain so little mineral matter as to be of no import; the only value, if any, being in the temperature. They are used principally for baths.

130. **Officinal Forms.—Aqua.** Natural water in the purest attainable state. **Aqua distillata.** Take of water 80 pints. Distill two pints, reject, then distill 64 pints. (U. S. P.) The transparency or color of distilled water should not be affected by lime water, H_2S, $BaCl_2$, $AgNO_3$, or $(NH_4)_2C_2O_4$.

The term **aqua,** in the U. S. P., is used to designate a solution of a gaseous or volatile body in water, as **aqua ammoniæ,**

aqua chlorinii. **Liquor,** a solution of a fixed or solid body, as **Liquor Ferri Nitratis, Liquor Plumbi Subacetatis.** A **Decoction** is a solution made with boiling water, usually of a vegetable product. An **Infusion** is a solution formed by subjecting a body for a short time to either cold or warm water. **Maceration** is the long continued action of water at the ordinary temperature. **Digestion,** the same with hot but not boiling water. **Lixiviation** or **leaching** is the process of pouring water upon a porous mass of any body, for the purpose of dissolving out any soluble matters.

131. **Physiological Use.**—Water exists in all the tissues of the body, and in all foods and drink. A healthy adult takes, on an average, about 2.5 litres of water in 24 hours; and loses by the skin, lungs, kidneys and fæces a little more than this; the **excess** coming from the oxidation of the hydrogen of the food and tissues. Water constitutes about from 65 to 70 per cent. of the whole body, being in slightly larger proportion in the young than in the adult body. The water in the tissues serves as a solvent for the various proximate principles intended for nourishment of the tissues, or coming from their waste, and intended for excretion. The evaporation from the skin serves to carry off the superfluous heat of the body.

132. **Hydrogen Peroxide** (H_2O_2).—The simplest way of preparing a diluted solution of this body is to pass a stream of CO_2 through water containing in suspension barium peroxide.

$$BaO_2 + CO_2 + H_2O = BaCO_3 + H_2O_2.$$

The insoluble barium carbonate may be separated by filtration, best through asbestos filters. This solution may then be evaporated in a partial vacuum. In its purest form, it is a syrupy, colorless liquid, having an odor resembling that of chlorine or ozone, and a tingling metallic taste. It is still liquid at —30° C. (—22°F.), but at temperatures above 22° C. (71.6°F.) it changes, rapidly into water and oxygen. This change takes place gradually at ordinary temperatures. In diluted solution it is much more stable, and may be boiled without suffering decomposition. The solutions of this substance are decomposed by many fine metallic powders. In most of its reactions it acts as an oxidizing agent. Argentic oxide, however, is reduced to the metallic state by it.

Tests.—To the suspected solution, add a little starch solution, then some potassium iodide, and finally a few drops of a solution of ferrous sulphate. If any hydric peroxide be present a blue color will appear. Very delicate.

Uses.—Oxygenated water, or hydric peroxide, is used as a bleaching agent for the hair and skin, converting brunettes into blondes. It is often used to renovate old pictures, the whites of which have become dingy. It has been used as a disinfecting solution, to ulcers, in ozæna, diphtheritic and scarlatinal sore throats, or where the membrane has invaded the nose. Also used as a test for pus in urine, with which it causes an effervescence.

THE CHLORINE GROUP.

	Symbol.	Atomic Weight.	State.
Fluorine,	F.	19	?
Chlorine,	Cl.	35.5	gas.
Bromine,	Br.	80	liquid.
Iodine,	I.	127	solid.

133. The elements of this group are electro-negative, fluorine being most negative, and the iodine least so. They have a characteristic, pungent odor, and act as disinfectants and bleaching agents. They enter into direct union with many of the metals to form binary compounds. They form compounds with hydrogen, with well marked acid properties. They have little affinity for oxygen, but form several oxacids and salts, all of which are rather unstable. They form the following compounds:—

HF.	—	—	—	—	—	—	—
HCl.	Cl_2O	Cl_2O_3	Cl_2O_4	HClO	$HClO_2$	$HClO_3$	$HClO_4$.
HBr.	—	—	—	HBrO	—	$HBrO_3$	$HBrO_4$.
HI.	—	—	I_2O_4	HIO	HIO_2	HIO_3	HIO_4.

FLUORINE.

F. — 19.

134. Little if anything is known of this element. It cannot well be prepared. The sources of fluorine compounds are the native fluor spar, calcium fluoride, and cryolite—a sodium and aluminum fluoride.

135. **Hydrogen Fluoride—Hydrofluoric Acid.**—This acid is obtained by the action of sulphuric acid upon powdered fluor spar, with the aid of a gentle heat.

$$CaF_2 + H_2SO_4 = CaSO_4 + 2HF.$$

The operation is usually conducted in a lead or platinum vessel, as the acid attacks glass and most metals. The acid is a colorless, transparent gas, with a pungent odor, very irritating to the skin and mucous membranes. It is readily soluble in water, forming a colorless, highly acid and corrosive liquid, with a pungent odor. Care must be taken, in using it, not to allow it to come in contact with the skin, as it produces a painful ulcer, which heals with difficulty, and also constitutional symptoms of considerable severity.

The boiling point of the liquid is between 15° and 20°C. (59 and 68°F.), and it has a specific gravity of 1.060. The most characteristic property of hydrofluoric acid, is its power of dissólving glass by removing its silicon. This property is utilized for etching glass. The article to be etched is first coated with a thin layer of melted wax or paraffin, and the characters are then scratched through the wax with a steel point, so as to expose the glass where the etching is to take place. If the liquid is to be used, a wall of wax is built up around the characters and the liquid is poured into the inclosure. The characters thus etched are transparent. It is more common to invert the glass, wax downward, upon a leaden dish containing the fluor spar and sulphuric acid, and expose it to the fumes until the etching is as deep as desired. The etchings in this case are opaque, presenting the appearance of ground glass, and more easily seen. Fluorine forms no oxides.

CHLORINE.

Cl. = 35.5.

D. = 35.5. Sp. Gr. = 2.47. Equivalence I, III, V or VII.

136. **Occurrence.**—Chlorine always occurs in combination, in nature. The chlorides of sodium, potassium, magnesium and calcium occur in springs. Usual source, sodium chloride or common salt.

137. **Preparation.**—By the action of sulphuric acid upon sodium chloride in the presence of manganic oxide.

$$2H_2SO_4 + MnO_2 + 2NaCl = \\ Na_2SO_4 + MnSO_4 + 2H_2O + Cl_2$$

Or, by acting upon manganic oxide with hydrochloric acid. $4HCl + MnO_2 = MnCl_2 + 2H_2O + Cl_2$.

For a slow continual evolution of chlorine, as for disinfecting purposes, moistened chloride of lime is exposed to the air. The calcium hypochlorite is decomposed by the carbonic dioxide of the air, and chlorine is set free. For a more rapid evolution, we may use the same salt with a dilute acid.

138. **Properties—Physical.**—At ordinary temperatures chlorine is a greenish-yellow, pungent, suffocating gas. It is

irrespirable, causing inflammation of the air passages. It is nearly two and a half times heavier than air; it is soluble in water, one volume of water dissolving nearly three volumes of the gas at 10° C. (50°F.) The solution, **aqua chlorinii** U. S. P., is a greenish-yellow liquid, possessing the properties of the gas, but slowly changing, in the light, into hydrochloric acid. It should bleach but not redden litmus paper. Under a pressure of four atmospheres, at ordinary temperatures, or at a temperature of —40° C. (—40° F.), it is condensed to a dark yellow liquid.

Chemical Properties.—The affinities of chlorine are very strong and extensive. It is characterized by its strong tendency to combine with hydrogen and the metals, to form chlorides. It combines directly with many elements (finely divided copper, antimony or arsenic), with the evolution of light and heat. Its attraction for hydrogen is so strong that a mixture of these gases combine with an explosion, when exposed to direct sunlight, the light of burning magnesium, or the electric light. It burns readily in an atmosphere of hydrogen. It is capable of existing in two **allotropic** states; the one active, and the other passive. The passive or inactive form is the one obtained when the gas is prepared in the dark; and when prepared in daylight it is very active in its properties. When the same element is capable of existing in two or more forms, having different properties, these forms are called **allotropic** conditions; the property is called **allotropism.**

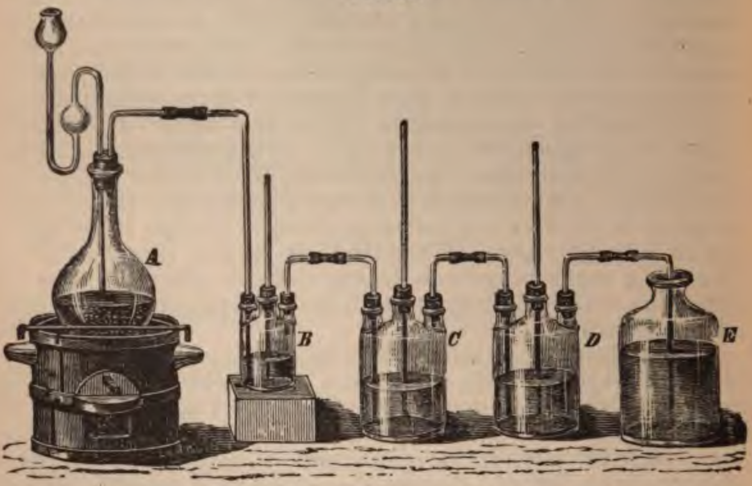
One of the most marked chemical properties of chlorine, is its affinity for hydrogen. So great is this affinity that many organic compounds are spontaneously decomposed by it; the chlorine combining with the hydrogen of the compound and setting the carbon free. A paper wet with turpentine and plunged into a jar of chlorine, takes fire and deposits the carbon as a dense smoke, while fumes of HCl fill the jar. The well known bleaching and disinfecting powers of chlorine are due to its affinity for hydrogen. Most vegetable colors, when moist, are readily discharged by chlorine. The chlorine combines with the hydrogen of the water and sets free the oxygen, which, in the nascent condition, is a powerful oxidizer, and decomposes the coloring agent or organized germ, as the case may be. In some cases the chlorine acts directly upon the organic matters, uniting with a portion of their hydrogen to form HCl, and a portion of it entering the molecule to take the place of the hydrogen removed. Thus, with marsh gas, hydrochloric acid and methyl chloride are produced. $CH_4 + Cl_2 = CH_3Cl + HCl$.

139. **Hydric Chloride.—Hydrochloric Acid.—Acidum Muriaticum—Acidum Hydrochloricum** (U. S. P.).—$HCl = 1 + 35.5 = 36.5$. Hydrochloric acid occurs very sparingly in nature. It is found in volcanic gases and in the gastric juice of mammals.

140. **Preparation.**—The acid is usually prepared from sodium chloride or common salt, by treatment with commercial sulphuric acid, with the aid of a gentle heat.

The process is sometimes conducted as a special process, but a large quantity of the acid is prepared, as a side product, in the preparation of sodium carbonate by Leblanc's process. The first step in this process is to treat the salt with sulphuric acid, and thus convert it into a sulphate. The hydrochloric acid set free by this process, is collected and sold as impure hydrochloric acid.

FIG. 32.



PREPARATION OF HCl.

$$H_2SO_4 + 2NaCl = 2HCl + Na_2SO_4.$$

The acid may be prepared in small quantity by the direct union of equal volumes of chlorine and hydrogen, under the influence of sunlight or the electric spark. (See Art. 85.)

141. **Properties.**—Hydrochloric acid is a colorless, transparent gas, having a pungent, penetrating odor, a sharp, sour

taste, an acid reaction, and produces great irritation of any tissue with which it comes in contact. It is irrespirable and extinguishes a flame. It is very soluble in water. One volume of this liquid dissolves 450 volumes of the gas at 15° C. (59° F.) This solution forms the ordinary **muriatic acid.** The specific gravity of the solution is 1.21, and contains about 40 per cent. of HCl. The sp. gr. of the gas (air $= 1$) is 1.264; the density (hydrogen $= 1$) is 18.25. Under a pressure of 40 atmospheres, at 10° C. (50° F.), it condenses into a colorless, limpid liquid, having a sp. gr. of 1.27. A strong solution in water fumes strongly in the air, giving off a part of the gas. On being heated, it gives off its acid rapidly. The strongest commercial acid contains from 25 to 30 per cent. of HCl. The commercial muriatic acid is yellow in color, due to the presence of ferric chloride. It also contains other impurities, and is used only for manufacturing purposes.

Acidum Muriaticum or **Acidum Hydrochloricum,** U. S. P., is a colorless liquid of sp. gr. 1.16. It contains 31.9 per cent. of HCl., with traces of impurities.

Acidum Muriaticum Dilutum, Acidum Hydrochloricum Dilutum, U. S. P., is made by diluting the stronger acid with water. (Strong acid 3 parts, distilled water 13 parts). The sp. gr. $= 1.049$, and contains 10 per cent. of HCl. Pure hydrochloric acid should be colorless, and when diluted with distilled water, should give no precipitate with H_2S, NH_4OH in excess, or $BaCl_2$, and should not dissolve gold leaf (absence of HNO_3).

142. **Chlorine and Oxygen.**—There are three oxides of chlorine known.

Cl_2O, **Hypochlorous Oxide** or **Anhydride.**
Cl_2O_3, **Chlorous Oxide** or **Anhydride.**
Cl_2O_4, **Chloric Tetroxide.**

The oxides of chlorine are all unstable compounds, very prone to decomposition, and of little importance.

143. **Hypochlorous Anhydride** is obtained by acting upon mercuric oxide with dry chlorine, as a blood-red mobile liquid below 20° C. (68° F.), $HgO + 2Cl_2 = HgCl_2 + Cl_2O$. Above this temperature it is a yellowish, pungent gas, resembling chlorine in many of its properties. It is a more powerful bleaching and disinfecting agent than chlorine, owing to the ease with which it decomposes. Water dissolves 200 times its volume of it, forming a colorless solution of **hypochlorous acid.** It sometimes decomposes with a slight jar, or even spontaneously, with the separation of chlorine and oxygen. Hypochlorous acid is un-

important, but it forms a series of salts called **hypochlorites.** A solution of the sodium salt, **liquor sodæ chlorinatæ** (U. S. P.), (Labarraque's solution), potassium hypochlorite, **liquor potassii chlorinatæ,** and calcium hypochlorite, or **chloride of lime,** are the most important compounds.

144. **Chlorous Anhydride** may be formed by treating potassium chlorate, $KClO_3$, with dilute nitric acid in the presence of arsenious oxide. It is a greenish-yellow explosive gas, soluble in water, with which it unites to form chlorous acid, $HClO_2$, a very unstable body which has not been isolated, but which forms **chlorites.** None of these salts are of importance.

145. **Chloric Tetroxide** may be obtained by treating potassium chlorate with strong sulphuric acid, as a yellow explosive gas. It is a powerful oxidizing agent. Below —20° C. (—4° F.) it is an orange-red liquid. It forms no corresponding acid.

146. **Chloric Acid** ($HClO_3$) and **perchloric acid** ($HClO_4$) are also known, but are very unstable, and like those before mentioned are powerful oxidizing agents. Potassium chlorate is the only salt of note, which will be described under potassiun.

BROMINE.

Br = 80.

Specific gravity 3.187 at 0° C. (32° F.). **Density of vapor** 80.

147. **History and Occurrence.**—Discovered by Balard, in sea salt, in 1826. It never occurs native, but is found combined with the alkaline metals and magnesium, in sea water, certain salt springs and the ashes of seaweeds. It is obtained from any of the above compounds by distilling them with manganic dioxide and sulphuric acid.

$$MgBr_2 + MnO_2 + 2H_2SO_4 = Br_2 + MgSO_4 + MnSO_4 + 2H_2O.$$

148. **Properties—Physical.**—Bromine is a blood-red fuming liquid, possessing a strong, disagreeable, pungent and irritating odor somewhat like that of chlorine. It corrodes animal tissues. It boils at 63° C. (145.4° F.), but gives off fumes at all temperatures above its freezing point, —24.5° C. (—12° F.). Water dissolves 3.2 per cent. of its volume at ordinary temperatures. It is soluble in alcohol, ether, chloroform and carbon disulphide, imparting its color to the solutions.

Chemical.—The chemical properties of bromine are similar to those of chlorine, but somewhat feebler. Bromine is poisonous. It may be recognized by its color, odor, or by the yellow or brown color of its solution in chloroform. It forms a yellow or orange color with starch paste. A solution of argentic nitrate precipitates it from its solutions, as a yellowish-white powder, which is soluble with difficulty in ammonium hydrate.

149. **Hydrogen Bromide** (HBr).—This acid may be prepared by treating phosphorus, immersed in cold water, with bromine, and distilling the resulting liquid. The bromine combines with the phosphorus, forming PBr_5, which is decomposed by the water into phosphoric and hydrobromic acids. It may also be prepared by the action of sulphuric acid (7 parts acid to 1 of water), upon a hot solution of potassium bromide. (Squibb.) It is a colorless gas, producing white clouds in the air, is readily soluble in water, forming a strongly acid solution, the properties of which closely resemble those of hydrochloric acid. The acid is used in medicine, and has about the same action as the other bromides, but is more active and rapid in its action.

150. **Bromine and Oxygen.**—No oxides of bromine are known. There are, however, three acids known, corresponding to those of chlorine. They are **hypobromous acid,** HBrO, **bromic acid,** $HBrO_3$, and **perbromic acid,** $HBrO_4$. They are of little importance, and but one salt of the three acids is of interest to the physician, and that is **sodic hypobromite,** NaBrO, used as a reagent for the estimation of urea. It is prepared in solution by adding bromine to a solution of sodium hydrate, having care to keep the mixture cool by immersion in cold water.

$$2NaOH + Br_2 = NaOBr + NaBr + H_2O$$

IODINE.

I = 127.

Specific gravity, 4.95.

151. **Occurrence.**—Iodine occurs in the same localities with chlorine and bromine, but in more sparing quantities than the latter. It is obtained mostly from the ashes of certain seaweeds collected in Scotland and France. The ash is leached with water, the solution evaporated down until the other salts crystallize out, leaving the more soluble iodides in solution. The solution is now treated with chlorine, to set the iodine free, aided by a sufficient heat to volatilize the iodine, which is condensed in suitable condensers and purified by sublimation.

152. **Properties.**—Iodine is a bluish-black crystalline solid, occurring in bright scales or tablets, which emit, even at ordinary

temperatures, a very irritating pungent vapor. When heated, they melt at 107° C. (225° F.), and boil at 175° C. (347° F.), giving a beautiful violet colored vapor, of the density of 127. This vapor condenses directly to the crystalline scales. When applied to the skin, it produces a yellowish-brown stain, which disappears after a time; but, after repeated applications, it causes irritation, thickening and desquamation of the cuticle. It is very slightly soluble in water (1 part to 7000), but the addition of potassium iodide renders it freely soluble. It is freely soluble in alcohol, ether, chloroform, glycerine, carbon disulphide and benzole. When pure it should leave no fixed residue when volatilized. **Cyanogen** is sometimes found as an impurity. This may be detected by heating the iodine on a water bath, covered with a cold watch glass, upon which the white crystals will condense. Its solutions, when cold, yield a deep blue color with boiled starch, which precipitates on standing. Boiled starch is used as an **antidote** in cases of poisoning.

153. **Medical Uses.**—It is used externally as a counter irritant and discutient; internally, as an antizymotic and alterative. In large doses, it acts as an irritant poison. It is eliminated by the kidneys, saliva and faucial mucous membrane, but not by the skin. In administering it, silver spoons should be avoided, as it attacks them. The following preparations of free iodine are officinal: **Iodum, Tinctura Iodinii,** a solution in alcohol (℥j to Oj). When freshly made, it is precipitated from this solution with water, but after some time it undergoes changes which prevent this. The so-called **colorless tincture** is made by adding ammonium hydrate to the above tincture, in sufficient quantity to decolorize it by converting the iodine into ammonium iodide.

Tinctura Iodinii Composita is a solution of iodine and potassium iodide in alcohol. **Liquor Iodinii Composita** (**Lugol's solution**), is a solution of iodine and potassium iodide in water. **Unguentum Iodinii** and **Unguent. Iodinii comp.,** the former containing iodine and the latter iodine and potassium iodide rubbed up with lard, are also officinal.

154. **Hydric Iodide, or Hydriodic Acid** (HI).—A solution of this acid is prepared by passing hydric sulphide through water containing iodine in suspension, until the iodine disappears, and then filtering from the precipitated sulphur. The acid, when pure, is a colorless gas, fuming in the air, having a penetrating odor resembling in most of its properties those of hydrochloric acid, although less stable and less active. **Acidum Hydrio-**

dicum Dilutum is officinal, but is prone to decomposition, with the liberation of iodine.

Iodides of sulphur, ammonium, arsenic, mercury, lead, iron, potassium and sodium are also used in medicine.

155. **Iodine and Oxygen.**—There are three or four unimportant and unstable oxides of iodine known. I_2O_3, I_2O_5, and possibly I_2O_7. Two oxygen acids are known, iodic, HIO_3 and periodic, HIO_4 or $H_2IO_6H_3$ (Thomsen & Basarow). The first is obtained by treating iodine with nitric acid, and the second by passing chlorine through an alkaline solution of sodium iodate.

The first, when pure, appears as a white crystalline solid, and the second as colorless crystals. Both are very soluble in water, are easily decomposed, and form salts.

The following compounds are also known : ICl, ICl_3, ICl_5, IBr, IFl_5 and NI_3. The last is a very explosive compound.

SULPHUR GROUP.

Sulphur, S = 32.
Selenium, Se = 79.5.
Tellurium, Te = 128.

156. **Group Properties.**—The elements of this group are all solid at ordinary temperatures. Their quantivalence is usually dyad, tetrad, or hexad. They combine with hydrogen to form hydracids, containing two atoms of hydrogen. In binary compounds they are usually dyad. They form two series of oxides and corresponding acids, RO_2 and RO_3, H_2RO_3 and H_2RO_4. In the first series their quantivalence is tetrad, and in the second hexad. Sulphur is the most electro-negative, and tellurium the least so, but all are decidedly negative.

SULPHUR.

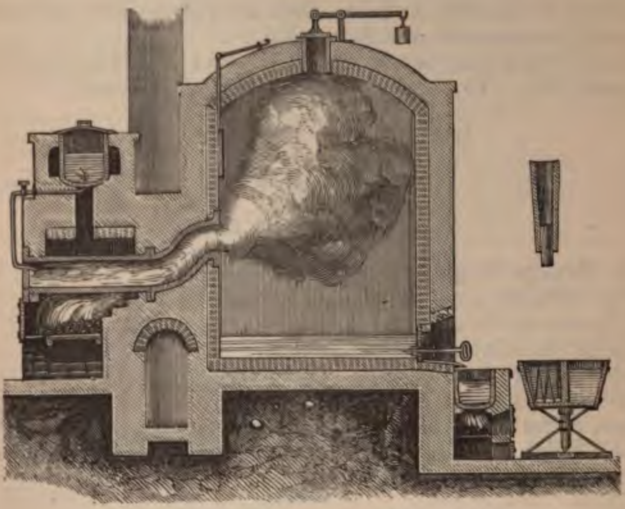
S = 32.

Sp. gr., 2.04. **Density of vapor**, 32. **Melts at** 114° C. (237° F.)

157. **Occurrence.**—Sulphur was known to the ancients. It occurs in volcanic regions, and is brought mostly from Sicily and Iceland. The sulphur occurs native, mixed with clay, from which it is separated by distillation. This element also occurs as sulphates in the minerals, gypsum, $CaSO_4 . 2Aq$, Barite, $BaSO_4$, etc.; and as sulphides of iron, copper, nickel, and, in fact, with many of the metals.

158. **Preparation.**—It is prepared from the native sulphur mixed with more or less earth, found in volcanic regions of Sicily, by distilling it from the non-volatile impurities. A second distillation is necessary to prepare the refined sulphur of the market. The second distillation is conducted in a retort, the vapor being conducted into a large chamber (Fig. 33), where, if the process is conducted slowly, it collects in the form of a crystalline powder called **flowers of sulphur.** If the process is conducted more rapidly, the chamber becomes hot, and the sulphur then condenses to a liquid at the bottom, whence it is drawn off into moulds forming ordinary **roll sulphur** or **brimstone.**

FIG. 33.



REFINING OF SULPHUR.

Sulphur is also obtained in some localities from iron pyrites, by piling it in heaps mixed with wood, to which fire is applied. The ore gives up a part of its sulphur, which melts and runs into cavities made under different parts of the heap.

159. **Properties—Physical.**—Sulphur, in its ordinary form, is a lemon-yellow solid, melting at 114° C. (234° F.), and boiling at about 440° C. (824° F.), giving off a brownish-yellow vapor, which in condensing returns directly to the solid state. Sulphur

is brittle, tasteless, odorless, a non-conductor of heat and electricity, and generates negative electricity when rubbed. It is insoluble in water, and almost so in alcohol. It is slightly soluble in aniline, phenol, benzol, benzine and chloroform. The best solvent is carbon disulphide, 100 parts of which dissolves 37 parts at ordinary temperatures.

Allotropic Forms.—Sulphur is capable of existing in three allotropic modifications; two crystalline, and one amorphous and plastic. The first variety is that found native, and occurs as octahedra of the second system. (See Art. 76.) It is freely soluble in carbon disulphide, from which the crystals separate on evaporation. The second variety is produced by crystallization from sulphur melted at high temperatures. This variety occurs as yellowish-brown, transparent prisms of the monoclinic system, of sp. gr. 1.98, insoluble in carbon disulphide, and gradually changes into the first variety. Since it crystallizes in two distinct systems, sulphur is said to be **dimorphous.** By heating sulphur, it melts at about 114° C. (234° F.) to a yellowish liquid; on raising the temperature to about 150° C. (302° F.), it becomes viscid and dark colored, and cannot be poured from the vessel; at a higher temperature, approaching its boiling point, it again becomes liquid. If sulphur, in this second liquid stage, be suddenly cooled by pouring into water, it assumes a soft, plastic, transparent mass, capable of being moulded like wax. This variety, like the preceding, gradually changes into the first variety, becoming opaque, yellow and crystalline.

Density of vapor at 500° C. (932° F.) $= 96 = S_6$. At 1000° C. (1832° F.) $= 32 = S_2$.

Chemical.—When heated in the air, sulphur takes fire and burns with a pale blue flame, and evolves abundant fumes of sulphurous anhydride, SO_2. It is generally strongly electro-negative and resembles oxygen in many of its compounds. In a few compounds it is electro-positive. It unites directly with many of the metals, especially when in the melted state, some metals taking fire and burning readily in its vapor. It forms the basis of a large and useful class of compounds, many of which resemble in composition the corresponding compounds of oxygen.

Thus, carbon disulphide, CS_2, corresponds to carbon dioxide, CO_2. Corresponding to hydric oxide, H_2O, we have hydric sulphide, H_2S, and to cyanic, CNOH, we have sulpho-cyanic acid, CNSH. Corresponding to carbonic acid, H_2CO_3, we have sulpho-carbonic acid, H_2CS_3.

160. **Uses.**—Sulphur is used in the arts, in the manufacture of sulphuric acid, H_2SO_4, as a bleaching agent for straw and woolen goods, and in the manufacture of matches and gunpowder. In medicine it is used as a parasiticide, and as a gentle laxative, although not as frequently as formerly. It is innocuous.

161. **Officinal Forms.**—1. **Sulphur sublimatum,** commercial flowers of sulphur. 2. **Sulphur lotum,** washed sulphur. Flowers of sulphur usually contain small quantities of sulphurous oxide, which is removed by washing with hot water. 3. **Sulphur præcipitatum, lac sulphuris,** milk of sulphur, is made by boiling for two hours sulphur and fresh slaked lime suspended in water.

$$3CaO_2H_2 + 2S_2 = 2CaS + CaS_2O_3 + 3H_2O.$$

The filtered solution of CaS and CaS_2O_3 is then diluted and treated with dilute hydrochloric acid as long as a precipitate forms.

$$2CaS + CaS_2O_3 + 6HCl = 3CaCl_2 + 3H_2O + 2S_2.$$

The calcium chloride, being very soluble, remains in solution. The sulphur is to be thoroughly washed with water, until free from acid. It is a fine white powder, easily suspended in water and viscid liquids. 4. **Unguentum sulphuris,** U. S. P., contains one part of sulphur to two of lard.

162. **Sulphur and Hydrogen.**—Two compounds of sulphur and hydrogen are well known—hydric sulphide, H_2S, and hydric persulphide, H_2S_2. The first only is of sufficient interest to merit description here. **Hydric sulphide, hydro-sulphuric acid** and **sulphuretted hydrogen** are synonymous terms. It is found in volcanic gases, in some mineral springs, and as a result of the decomposition of organic matter containing sulphur. At an elevated temperature, the two elements may be made to unite directly.

163. **Preparation.**—The usual method of preparing it is to act upon a sulphide, usually ferrous sulphide (FeS), with dilute sulphuric acid.

164. **Properties.**—H_2S is a colorless, transparent gas, of an unpleasant odor resembling that of rotten eggs, soluble in water, to which it imparts acid properties. It is somewhat heavier than air, its density being 17 and its sp. gr. 1.177. At a temperature of —74° C. (—101.2° F.), or under a pressure of 17 atmospheres at 10° C. (50° F.), it condenses to a colorless, mobile liquid, which at —85° C. (—121° F.) becomes an ice-like solid.

It burns with a blue flame, producing water and sulphurous oxide or anhydride, SO_2.

$$2H_2S + 3O_2 = 2H_2O + 2SO_2.$$

If the supply of oxygen is deficient, H_2O is produced, while the sulphur is deposited free. It is decomposed by chlorine, bromine, iodine, and oxidizing agents in general. It is also decomposed by sulphurous oxide.

When the gas is allowed to bubble through a solution of an alkaline hydrate, the sulphur and oxygen exchange places, with the formation of a sulphydrate. $KOH + H_2S = KSH + H_2O$. When it is passed through a solution of a metallic salt, it forms, in many instances, a sulphide of the metal. $CuSO_4 + H_2S = CuS + H_2SO_4$. It is, on this account, largely used in the laboratory as a reagent for the separation of the metals from one another. Minute quantities of H_2S may be detected by its odor, or by the brown or black color it imparts to a paper moistened with a solution of plumbic acetate. Its principal use is as a reagent.

165. **Physiological.**—When inhaled, it is not an irritant, but a narcotic poison, even when largely diluted with air. According to Faraday, birds die in air containing $\frac{1}{1500}$ of it, and dogs in one containing $\frac{1}{800}$. According to Letheby, human beings cannot live in an atmosphere containing more than one per cent. Its action is principally a reducing one upon the hæmoglobin of the blood, and prevents this fluid from absorbing oxygen, although it probably does not combine with it. (Wurtz.) Hydrosulphuric acid is formed in the intestine, from the decomposition of albuminous matters, especially where there is any impediment to digestion, or to the onward movement of their contents. It also sometimes occurs in abscesses, in the urine and bladder.

This gas is almost a constant ingredient in the gas of sewers and privy vaults, existing free or combined with ammonium as ammonium sulphydrate. Poisoning by this gas may be acute or chronic. The latter is more common, producing a febrile state, with malaise and general debility. The injurious effects of sewer gas are partially due to this form of poisoning, although not entirely. Occasionally this gas is so concentrated in sewers that those who enter them suffer with acute poisoning, fall almost instantly, and if not rescued, die in a short time. The treatment, in such cases, should consist in pure air, or oxygen, with brandy and water. Chlorine water, or a mixture of potassium chlorate and dilute hydrochloric acid, may be administered interganic matter gives to the acid a dark color. Arsenic and iron are sometimes found in the commercial acid.

The C. P. acid should not contain lead, iron or organic matter. It often contains oxides of nitrogen and sulphurous acid. Sulphuric acid and soluble sulphates are easily detected by the use of barium chloride, or nitrate, with which they form a white precipitate insoluble in hydrochloric acid. Lead, silver and mercury must first be removed with H_2S, and the solution diluted.

176. **Fuming Sulphuric Acid—Nordhausen Acid** ($H_2S_2O_7$), is obtained by distilling ferrous sulphate. The first portions are a white crystalline solid, fusing at 35° C. (95° F.), having the above composition.

The commercial acid is a brown, oily liquid, fuming in the air, and hissing when dropped into water. Some chemists regard it as a solution of SO_3 in H_2SO_4. When heated, it gives off SO_3 and H_2SO_4.

$$H_2S_2O_7 = H_2SO_4 + SO_3.$$

It is used in manufacturing alizarine, eosin, etc., and as a solvent of indigo. It forms a series of salts called disulphates.

177. **Sulphur with Chlorine, Bromine and Iodine—Sulphurous Chloride** (S_2Cl_2).—A yellow, volatile, fuming liquid, formed by distilling sulphur in an atmosphere of chlorine gas, and having a powerful solvent power for sulphur and sulphuric chloride.

It is decomposed by water, but mixes with benzol and carbon disulphide. SCl_2, and several oxychlorides, are known. **Bromine** unites directly with sulphur to form a red, unstable liquid, probably consisting mostly of S_2Br_2.

Iodine and Sulphur combine directly when gently heated, even under water. When 127 parts of iodine and 32 parts of sulphur are heated they form a steel-gray crystalline mass, S_2I_2, **sulphuris iodidum** (U. S. P. and B. P.), said to be a powerful remedy in certain skin diseases. It melts at 60° C. (140° F.), and is insoluble in water. Other iodides have been described, but are unimportant.

SELENIUM AND TELLURIUM.

Se. 79.5. Te. 128.

elements are rare and of no special interest to the

NITROGEN GROUP.

Nitrogen, N = 14 III or V.
Phosphorus, P = 31 III or V.
Arsenic, As = 75 III or V.
Antimony, Sb = 122 III or V.
Bismuth, Bi = 210 III or V.

179. **Group Characteristics.**—A well defined group with nitrogen at the negative end and bismuth at the positive. The atomic weights form a graded series from 14 to 210. The first is a gas; the second a volatile solid; the third a volatile, crystalline, metallic-looking body, showing a slight tendency to alloy with metals and combine with acids; the fourth less easily volatilized, crystalline, possessing a brilliant lustre, alloying with metals, and showing a tendency to act the positive rôle with acids; the fifth, also crystalline, having a metallic lustre, and showing more marked positive tendencies. They are all both triad and pentad, and form two series of compounds.

The following will exhibit the relations of some of the most important compounds:—

Hydrides.	Chlorides.	Oxides.	Sulphides.
NH_3	NCl_3 ...	N_2O_3, N_2O_5	... ...
PH_3	PCl_3, PCl_5	P_2O_3, P_2O_5	P_2S_3, P_2S_5
AsH_3	$AsCl_3$, $AsCl_5$	As_2O_3, As_2O_5	As_2S_3, As_2S_5
SbH_3	$SbCl_3$, $SbCl_5$	Sb_2O_3, Sb_2O_5	Sb_2S_3, Sb_2S_5
...	$BiCl_3$, ...	Bi_2O_3, Bi_2O_5	... ...

NITROGEN.

Symbol, N. At. wt. 14. Equivalence, I, III or V. **Density, 14.**
Wt. of 1 litre, 1.256 grms. Sp. gr., (Air = 1) **0.971.**

180. **Occurrence.**—Exists free in air, mixed with oxygen. It is also found free in the gases found in the stomach, large and small intestine, the urine, etc. Combined, it occurs as nitrates of potassium, sodium and calcium, in ammonia, and in many vegetable and animal bodies of the proteid group.

181. **Preparation.**—From the air, by burning phosphorus in a confined space until the oxygen is removed, or by passing air over copper or iron turnings heated to redness, the nitrogen prepared by both these methods contains small quantities of other gases found in the air. To prepare it pure, heat ammonium nitrite (NH_4NO_2).

182. **Properties—Physical.**—A colorless, transparent, odorless, tasteless, incombustible gas, not a supporter of combustion or

of animal respiration. It is not poisonous; very sparingly soluble in water or alcohol. One part of water dissolves, at the ordinary temperature and pressure .025 parts of this gas. Chemically, nitrogen is characterized by its inertness. It unites directly with magnesium, boron, vanadium and titanium. Indirectly, it forms a great number and variety of compounds, many of which are unstable. Under the influence of electric discharges nitrogen can be caused to unite with hydrogen to form ammonia, NH_3, and with oxygen to form nitrous and nitric oxides. From this source most of the nitrogenous products necessary to sustain plant life are primarily derived.

THE ATMOSPHERE.

183. The **atmosphere** is composed principally of nitrogen and oxygen mixed together in the proportion of 20.93 parts of oxygen by volume to 79.07 parts of nitrogen, and, by weight, 23 parts of oxygen to 77 parts of nitrogen.

Although air is a mixture and not a definite compound, it is remarkably constant in composition. Regnault found in 233 analyses of air, at different times and places, that the per cent. of oxygen by volume varied between 20.908 and 20.999. That air is a mixture is proved by: 1st, its gases are not present in the proportion of their atomic weights; 2d, air can be made, to answer all the properties of the atmosphere, by a mechanical mixture of the gases; 3d, solvents for oxygen, as an alkaline solution of pyrogallic acid, remove this gas from the air; 4th, each gas dissolves in water independently of the other, and with its own solubility; thus, by expelling the air from water, by boiling, and analyzing it, we find it to correspond to that calculated from the known solubility of the two gases.

The analysis of the air expelled from water contains 33 per cent. of oxygen, and 67 per cent. of nitrogen; it is, therefore, much richer in oxygen than the atmosphere. Owing to the rapid diffusion of the gases, the disturbances in composition due to the respiration of animals and manufacturing processes, are soon restored. Besides the two chief gases found in the air, there are various other ingredients found in small quantities; as water vapor, carbon dioxide, ozone, ammonia, nitric and nitrous acids, hydrocarbons, solid particles of dust, sodium chloride, vegetable germs or spores, bacteria, etc. Air in which animals are confined, also contains some of the organic exhalations from their bodies. In the neighborhood of large cities, various other substances are poured into the air from manufacturing establishments.

The essential ingredients are oxygen, nitrogen, carbon dioxide, and watery vapor. The rest of those enumerated, may be regarded as accidental, and not essential to the growth of plants and animals.

184. **Watery Vapor.**—The proportion of watery vapor in the air varies considerably with the temperature and locality. The air is seldom saturated in the daytime, and contains less inland than near large bodies of water. The higher the temperature of the air the more moisture it will hold; thus, at 0°C. (32°F.), 1 cu. metre (1.3 cubic yards) is saturated by 5.4 grammes (83.3 grs.) of water, and at 25°C. (77°F.), the ordinary temperature, it requires 22.5 grammes (347 grs.). Or, at 77°F., one cubic yard will be saturated by 267 grains of water, and one cubic foot by about 10 grs. In reality, the air will seldom be found to contain more than 60 or 70 per cent. of this amount. When one cubic metre (1.3 cu. yds.) of an atmosphere, saturated at 25°C. (77°F.), is cooled down to 0°C. (32F.), it will deposit as dew, rain, or frost, $22.5 - 5.4 = 17.1$ gms. (263.8 grs.). The temperature at which air begins to deposit its moisture, on being cooled, is called the **dew point.** The dew point will depend upon the amount of water actually present in the air. The amount of moisture is determined by passing a known volume of air through tubes containing calcium chloride, which absorbs the water. The increase in the weight of the tube gives the weight of water. The amount of vapor in air varies from .3 to 1.6 per cent. by volume. The dampness of the air does not depend upon the amount of water it contains, but upon the degree of saturation. A cold damp air, when heated, becomes dry; hence, the necessity of supplying moisture to the heated air in our rooms in winter. A very dry air irritates the air passages, produces dryness of the skin and malaise. A very moist atmosphere checks evaporation from the skin and lungs, raises the bodily temperature, and soon becomes oppressive.

185. **Carbon Dioxide.**—The average amount of carbon dioxide, CO_2, in country air is 4 parts in 10,000, and varies from 3 to 6 parts. It is greatest near large cities and manufactories, greater during the night than the day, on land, and the reverse on the ocean. Plants remove it from the air in the daytime, and the cooler water at night, more than the warmer water during the day. (See Carbon dioxide.)

186. **Ammonia.**—This exists in the air, in very minute quantities, in the form of carbonate, nitrate and nitrite, the result of the decomposition of animal and vegetable organic matters.

It is especially evolved from urinals, privy vaults and horse stables. It is washed out of the air by falling rain, and is taken up from the soil by plants.

187. **Nitric and Nitrous Acids** occur in extremely minute quantities, and are produced by the direct union of oxygen and nitrogen in the presence of watery vapor, under the influence of discharges of lightning. They exist principally in combination with ammonia.

Hydrocarbons, the principal of which is marsh gas, are frequently found in the air of cities, coal mines, wells and swampy districts. It is produced by the decomposition of vegetable matter under water, and in some industrial processes.

188. **Accidental Gases in the Air.**—The gases generated in certain manufactures are sometimes allowed to escape into the air from the rooms in which workmen are employed. Some of these are harmless and others hurtful. Among the first class may be mentioned **carbon dioxide,** when not in too large quantities, and **ammonia.** To the second class belong **hydric sulphide, ammonium sulphydrate, sulphurous oxide** in large quantities, vapors of **mineral acids, carbon disulphide,** etc.

Hydric Sulphide, or **Sulphuretted Hydrogen,** is found in certain tunnels and mines, caused by the decomposition of iron pyrites. It is also found in the air of some marshes and sewers.

The symptoms produced by breathing small quantities of this gas, are those of debility and anæmia; in larger quantities, headache, vertigo, weak pulse, sweating and prostration.

Sulphydrate of Ammonia produces nearly the same symptoms as hydric sulphide. It occurs in the air of sewers and privy vaults. Both these substances are easily destroyed by chlorine or sulphurous oxide.

Sulphurous Oxide, unless in considerable quantities, and in a closed room, does not seem to have any deleterious effect upon the workmen. In bleachers it sometimes produces irritation of the bronchial tubes.

Hydrochloric Acid, Nitric Acid, and **Chlorine** in considerable quantities are very irritating to the lungs and conjunctiva. **Carbon Disulphide** produces unpleasant and deleterious effects upon workmen exposed to air containing it; as headache, giddiness, nervous depression and loss of appetite.

189. **Suspended Matters.**—A great variety of solid particles, or dust, are found in the air at all times. These consist of fragments of wood, textile fabrics, metals, etc., pollen of plants,

bacteria germs, etc. These suspended particles may be regarded as impurities, and many of them are injurious to health. Workmen in various trades are seriously affected by the dust to which they are exposed; as miners, especially of lead and coal, grinders of metals, wool sorters, rag pickers, feather dressers, etc. The irritation of the dust of these and other trades may cause chronic bronchitis, emphysema, phthisis, or chronic poisoning. Germs of various kinds are believed to cause many of the contagious and malarial diseases, and may be carried some distance in the air. Some of these germs seem to be easily oxidized, while others are very persistent. The best disinfectants for their destruction, are free ventilation and consequent dilution, chlorine, bromine, iodine and sulphurous oxide.

190. **Disinfectants, Germicides,* Antiseptics, Deodorizers.**—The presence of odors and organized "germs" in the air, often require the use of one of the above agents. **Disinfectants** are a class of bodies which are supposed to destroy the germs, and thus prevent them from causing their specific action either upon the human body or in decomposable organic bodies or solutions.

The most efficient of these is heat. Organized germs may be filtered from the air by passing it through cotton wool; or they may be removed by inclosing the air in an air-tight box or chamber, the insides of which are moistened with glycerine. (Tyndal.) Ozone, chlorine, bromine, iodine, sulphurous oxide, mercuric, zinc, aluminium, magnesium and calcium chlorides, potassium chlorate, potassium permanganate, carbolic, boric, cresylic and sulphuric acids, thymol, menthol, camphor, etc., are among the disinfectants most used.

Antiseptics are agents which retard or entirely prevent putrefaction or growth of microscopic germs and organisms. While **disinfectants** destroy the cause of infection, **antiseptics** prevent the development of these causes. Low temperature retards putrefaction, and is, therefore, an antiseptic agent. These two terms are frequently used interchangeably.

Asepsis is a condition of entire absence of any germs or cause of infection. **Deodorizers** are bodies used to destroy offensive odors. They may be either solid, liquid or gaseous. **Solids**—dry earth, lime, charcoal, ferrous sulphate, carbolates of calcium, sodium and magnesium. **Liquids**—solutions of plumbic nitrate

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* A *germicide* is an agent which has the power of killing the germs, and thus preventing their growth. A disinfectant destroys the infectious properties of a septic matter, whether this be due to germs or some other agent.

(Ledoyen's fluid), zinc chloride (Burnett's fluid), potassium or sodium permanganate (Condy's fluid), a mixture of copper and zinc sulphates (Lanande's disinfectant), solutions of ferric chloride, of ferrous sulphate, hypochlorites, etc., are among the best known. **Gases**—pure air, ozone, chlorine, bromine, and sulphurous oxide are those most effective. Fumigations with tar, herbs, and various aromatic substances, only disguise the offensive odors, but do not destroy them.

The ordinary offensive odors are due to hydric sulphide (H_2S), ammonium sulphhydrate (NH_4HS), phosphoric hydride (PH_3), and complex ammonium compounds. Chlorine, ozone, and nitrous oxides will destroy these gases by oxidation, and thus destroy the odor.

It should be remembered that these odors, in themselves, may not be in any degree injurious to health, when in small quantity, but they serve to warn us of the presence of other products of putrefaction which accompany them, and which are injurious. The fact that efficient disinfection of the air can prevent the spread of the contagion of disease is well known. Chlorine and sulphurous oxide are the two agents most in use, and of these the former is very much to be preferred, but the latter is used for furnished rooms, because of its less destructive action on articles exposed to it.

It is doubtful whether organized germs can be destroyed in the air by any disinfectants, except in tightly closed rooms. The attempt to disinfect the air of rooms with the various so-called "disinfectants" of the market is worse than useless. It engenders a feeling of security where there is none. These floating germs can certainly stand as much, and in most cases, more than man, and therefore, no room can be disinfected while it is occupied by human beings. The author has found by experiment, that most of the ordinary antiseptics, when diffused through the air of an ordinary room, are almost without action on putrefactive bacteria, unless the quantity be great enough to make the air irrespirable.

The following table shows the amount of water it is necessary to add to one part of the substance named, which barely permits the development of bacteria in meat infusions, according to M. Jalan de la Croix:—

1 Part.	Water, Parts.	1 Part.	Water, Parts.
Alcohol	30	Oil of mustard	5,734
Chloroform	134	Sulphurous acid	7,534
Borax	107	Aluminium acetate	7,535
Eucalyptol	308	Salicylic acid	7,677
Phenol (Carbolic acid)	1,002	Mercuric chloride	8,358
Thymol	2,229	Calcium hypochlorite	13,092
Potass. permanganate	3,041	Sulphuric acid	16,782
Picric acid	3,041	Iodine	20,020
Borated sodium salicylate	3,377	Bromine	20,875
Benzoic acid	4,020	Chlorine	34,509

[illegible] says of iodine, that 1 part to 12,000 destroys the con- [illegible] charbon, and 1 to 10,000 of septic blood. Billroth says [illegible] chloride 1 to 20,000, thymol and benzoate of sodium 1 [illegible] and benzoic acid and creasote 1 to 1000, prevent the de- [illegible] of bacteria. Koch says of mercuric chloride, 1 to [illegible] kills most micro-organisms, and 1 to 1000 destroys resting [illegible]

The results of different experimenters are so widely discordant, [illegible] make no attempt to reconcile them. The following table [illegible] the results of experiments upon commercial disin- [illegible]. The first column gives the per cent. of the agent ne- [illegible] to kill anthrax and bacillus subtilis. The second [illegible] the per cent. of the agent which failed to produce this result.

LIST OF COMMERCIAL DISINFECTANTS (*Sternberg*).

NAME.	Per cent. in, when active in two hours.	Per cent. in, which failed in two hours.
Little's Soluble Phenyle	2	1
Labarraque's Solution (U. S. P.)	7	5
Liquor Zinci Chloridi (Squibb's)	10	7
Feuchtwanger's Disinfectant	10	8
Labarraque's Sol. (Frere, Paris)	15	10
Phenol Sodique	15	10
Platt's Chlorides	20	15
Girondin Disinfectant	25	15
Williamson's Sanitary Fluid	25	20
Bromo-chloralum	25	20
Blackman's Disinfectant	30	20
Squibb's Impure Carbolic Acid	—	50
Buchan's Disinfectant	—	50
Phenol Sodique (Paris)	—	50
[illegible]	—	50
Hypochlorite of Sodium or Calcium. Available Chlorine	0.25	...

NITROGEN AND HYDROGEN.

Ammonia. $NH_3 = 14 + 3 = 17$.
Density = 8.5. Sp. gr. .59.

[illegible] from the decomposition of animal or vegetable [illegible], either spontaneously, or by the aid of [illegible] by distilling camels' dung, in Libya, near [illegible] Ammon. When horns, clippings of hides, [illegible] closed retorts, ammonia is given off. The [illegible] present, is from the ammoniacal liquors [illegible] contains about 2 per cent. of nitrogen, which is mostly given off as ammonia. The ammonia liquor is treated with hydrochloric acid, and evaporated to dryness, when an impure ammonium chloride, **sal-ammoniac**, is obtained. This may be purified by recrystallization or sublimation. This salt, heated with lime (CaO) gives off its ammonia. This is conducted through a series of Woulfe bottles containing water, in which the gas dissolves, forming aqua ammonia, from which the other compounds may be prepared.

$$2NH_4Cl + CaO = 2NH_3 + H_2O + CaCl_2.$$

192. **Properties.**—A colorless, transparent, pungent, irrespirable gas. Does not support combustion or burn in air, but burns with difficulty in an atmosphere of oxygen, forming water and free nitrogen. It has a strong alkaline reaction on moistened litmus paper, which, however, is not permanent, owing to the volatility of the ammonia.

It is lighter than air. Liquefies at —40°C. (—40°F.); or at 10°C. (50°F.), under a pressure of 6.5 atmospheres, to a colorless liquid of specific gravity 0.76, which solidifies at —75° C. (—103°F.).

It is very soluble in water; one volume of water at 15°C. (59°F.) dissolves 783 volumes of the gas with the evolution of heat, forming the solution known as **aqua ammoniæ** or **liquor ammoniæ**, which may be regarded as a solution of ammonium hydrate, NH_4–O–H.

This solution, on being heated, gives up most of the gas again. **Aqua ammoniæ fortuor** (U. S. P.) contains 28 per cent., by weight, of the gas, and has a sp. gr. of 0.900 at 15°C. (59°F.). **Aqua ammoniæ** (U.S.P.) contains 10 per cent. by weight, and has a sp. gr. 0.959 at 15°C. (59°F.) It is a colorless, transparent liquid, with a pungent odor, and alkaline taste and reaction. It forms, by direct union with the acids, a series of salts containing the compound radical NH_4, called ammonium. Ammonia is volatile, and hence it is sometimes known as the **volatile alkali.** The compounds of NH_4 closely resemble those of Na and K, and will be considered with them. The strong solutions of the gas act as a caustic upon animal tissues, and are, therefore, corrosive poisons.

193. **Composition.**—This may be determined by decomposing the gas by passing a series of electric sparks through a quantity of it inclosed in a eudiometer tube over mercury. The volume increases until double the original volume is reached. By introducing a quantity of oxygen equal to that of the am-

monia used, and igniting the gases by the same spark, the hydrogen and oxygen combine, and after condensing leave the nitrogen, which occupies one-half the original volume, or one-fourth the volume of the mixed hydrogen and nitrogen after the decomposition. It is thus shown to be composed of one-fourth nitrogen and three-fourths hydrogen.

We may also arrive at the same result in the following manner:—

Prepare a glass tube of about 1 c.c. ($\frac{1}{3}$ in.) calibre, closed at one end; through the stopper in the open end pass a funnel tube drawn to a point and provided with a stop-cock. Fill the tube with pure dry chlorine, and insert the cork. Fill the funnel tube with strong ammonium hydrate solution, open the stop-cock, and allow a portion of the liquid to enter the tube. The chlorine decomposes the ammonia gas, combining with its own volume of hydrogen and setting free the nitrogen in combination with it. By removing the stopper under water, the water will rise to fill the tube, excepting that portion occupied by the nitrogen, which will be found to be one-third of the whole tube.

Now, as the chlorine combined with its own volume of hydrogen, or with the tube full, and left one-third of that volume of nitrogen, it is easy to see that the ammonia was composed of three parts by volume of hydrogen, and one part of nitrogen. Since gaseous molecules all occupy the same space, three molecules of hydrogen and one of nitrogen form two of ammonia: $3H_2 + N_2 = 2NH_3$.

The compounds of ammonium with acids will be considered under the head of salts of alkaline metals.

194. **Tests.**—Smell. Fumes with HCl. Moistened red litmus paper is changed to blue by it. Nessler's test, see Art. 125.

195. **Hydroxyl Amine** (NH_3O or $N\begin{smallmatrix}H\\H\\O-H\end{smallmatrix}$).—This compound, closely related to ammonia, may be regarded as a molecule of ammonia in which one hydrogen atom has been replaced by the hydroxyl radical. Prepared by treating tin with dilute nitric acid, or a mixture of this acid and hydrochloric acid. The nascent hydrogen generated, reduces the acid and forms the above compound, which combines directly with the remaining free acid.

$$HNO_3 + 3H_2 = H_3NO + 2H_2O$$
$$\underset{\text{Hydroxyl Amine.}}{H_3NO} + HCl = \underset{\text{Hydrochloride of Hydroxyl Amine.}}{NH_2OH.\ HCl}$$

It is an unstable liquid, not obtainable in a free state, and possesses decided basic properties—blues reddened litmus paper, and combines directly with acids to form salts.

196. **Nitrogen Chloride** (NCl_3).—When chlorine, in excess, is made to act upon ammonia, or a solution of ammonium chloride, the chlorine at first sets free nitrogen, and forms ammonium chloride; the excess of chlorine then acts upon the ammonium chloride, to form nitrogen chloride.

$$NH_4Cl + 3Cl_2 = NCl_3 + 4HCl.$$

197. **Properties.**—A yellow, oily liquid, insoluble in water, possessing a disagreeable, irritating odor. Sp. gr. = 1.653. It is very explosive, and in contact with any combustible matter, explodes spontaneously. It should not be prepared in large quantity.

198. **Nitrogen Iodide** (NHI_2 or NI_3). **Preparation.**—By lightly triturating iodine, in a mortar, with strong ammonium hydrate, or by pouring an alcoholic solution of iodine into strong ammonia water.

199. **Properties.**—A brownish-black solid, insoluble in water, which, when spread out on filter paper and dried, explodes with the slightest touch, or by a gentle breeze; the explosion, however, is not nearly so violent as that of the chlorine compound.

200. **Nitrogen** and **Oxygen.**—Five oxides of nitrogen are known, whose names, graphic formulæ and corresponding acids are as follows:—

Nitrous Oxide	$N-O-N$	**Hyponitrous Acid**	$H-O-N$
Nitric Oxide	$O=N--N=O$		
Nitrous Anhydride	$O=N-O-N=O$	**Nitrous Acid**	$H-O-N=O$
Nitric Peroxide or Tetroxide	${}^{O=}_{O=}N-O-N=O$		
Nitric Pentoxide or Anhydride	${}^{O=}_{O=}N-O-N{}^{=O}_{=O}$	**Nitric Acid**	$H-O-N{}^{=O}_{=O}$

201. **Hyponitrous Oxide. Nitrous Oxide, Laughing-gas. Nitrogen Monoxide. Nitrogen Protoxide** (N_2O).—Discovered in 1776, by Priestly. Anæsthetic effect first discovered by Sir Humphrey Davy. First used in dentistry by Wells, of Hartford, Ct. First came into notice as an anæsthetic in 1863.

202. **Preparation.**—By gently heating ammonium nitrate, when it decomposes into hyponitrous oxide and water.

$$NH_4NO_3 = N_2O + 2H_2O.$$

For anæsthetic purposes, care should be exercised to keep the temperature of the retort between 210°C. (410°F.) and 250°C. (482°F.), as below the former the decomposition does not take place, but the salt sublimes; while above the latter, nitrogen dioxide and trioxide are generated. As an additional safeguard, the gas should be caused to bubble through solutions of sodium hydrate and ferrous sulphate, to remove these higher oxides.

203. **Properties.**—A colorless, odorless, sweetish-tasting gas, slightly soluble in water, more so in alcohol. Density 22. Sp. gr. 1.527. Under a pressure of 50 atmospheres at 7°C. (45°F.) it condenses to a colorless liquid, which resumes the gaseous state as soon as the pressure is removed, the temperature sinking so low as to freeze a portion of the liquid into a white, snow-like solid. Sp. gr. of liquid 0.908. Boiling point —88°C. (—126°F.) Freezing point about —101°C. (—150°F.) It is neutral in reaction, *i. e.*, neither acid nor alkaline. It supports the combustion of bodies, very much like oxygen; this is due to the fact that the heat of the burning bodies decomposes the gas, giving an atmosphere about them containing twice as much oxygen as ordinary air. For anæsthetic purposes, the liquefied gas is now sold in wrought-iron cylinders, provided with a stopcock, so that the gas can be drawn from the cylinder as needed.

204. **Physiological Effects.**—Nitrous oxide causes, when first inhaled, an exhilaration, then anæsthesia, and finally, asphyxia. It will not support the respiration of plants or animals. It seems to act partly by excluding air, and partly by its direct effect upon the nervous system, but does not enter into any chemical combination in the blood; it simply dissolves in this fluid. When mixed with oxygen, and administered under an increased pressure, the anæsthesia may be kept up for a long time with safety. Deaths from its inhalation are rare. It does not undergo decomposition in the blood. It is much used for short operations, and especially for the extraction of teeth, opening abscesses, felons, etc. Recovery is prompt and complete within a few minutes after its withdrawal. A solution in water, containing five volumes of the gas, has been administered internally.

205. **Hyponitrous Acid** (HNO).—This acid may be prepared by the action of hydrochloric acid on the silver salt.

The potassium salt (KON) is formed by the action of sodium amalgam on potassium nitrite or nitrate; preferably the former: $KONO_2 + 2H_2 = KON + 2H_2O$.

The silver salt is a yellow, almost insoluble powder.

206. **Nitric Oxide or Nitrogen Dioxide.** NO or N_2O_2.—Prepared by the action of nitric acid upon copper.

$$3Cu + 8HNO_3 = 3Cu(NO_3)_2 + N_2O_2 + 4H_2O.$$

207. **Properties.**—A colorless, transparent gas, very sparingly luble in water, more soluble in alcohol. Density 15, sp. gr. 039.

The density would make the molecular weight 30, and the formula NO, which is anomalous, as in this case nitrogen must be considered as a dyad. The ordinary laws of valence would make it N_2O_2. It is probable that at lower temperatures this is the proper formula, and at the higher temperature dissociation takes place; N_2O_2 splitting up into NO, NO, as has been proven to occur in the case of N_2O_4.

By cold and pressure, the gas has been reduced to a liquid. Bodies which evolve considerable heat in burning, as, for example, phosphorus, burn in this gas; first decomposing it, and then uniting with its oxygen. In contact with free oxygen, or air, it takes up this gas and is converted into N_2O_4, or N_2O_3, according to the amount of oxygen present. In both cases it gives a reddish-brown colored gas. A **test** for free oxygen. It is rapidly absorbed by a solution of ferrous sulphate, to which it imparts a deep brown color. Its action on the economy is not known. It forms no corresponding acid.

208. **Nitrous Anhydride** (N_2O_3).—Prepared by the direct union of nitric oxide (NO) and oxygen, mixed in the proportion of four of the former to one of the latter. Also, by warming nitric acid with starch, or arsenious acid, and by the action of the peroxide on cold water.

$$2N_2O_4 + H_2O = N_2O_3 + HNO_3.$$

209. **Properties.**—A dark blue liquid, boiling at 0°C. (32°F.), with partial decomposition into NO and N_2O_4 which recombine on cooling. It combines directly with water, producing nitrous acid (HNO_2), which, on warming, decomposes into nitric acid and nitric oxide.

$$3HNO_2 = HNO_3 + 2NO + H_2O.$$

As will be seen from the above, this oxide is very unstable.

210. **Nitrous Acids and Nitrites.**—The acid is not known in a pure state, but several of its salts are known. The nitrites are formed by heating the nitrates, when they give off a part of their oxygen. The action is rendered easier, if lead or some other oxidizable metal be added to the fusion. The nitrites are produced in nature, by the oxidation of nitrogenous organic matter, accompanied by certain forms of microscopic life. Such nitrification takes place in waters polluted with organic matter, and normally in the soil. The acid then combines with bases found in the water or soil. The presence of nitrites in water, is, for this reason, looked upon as an evidence of previous contamination with nitrogenous matter. Further oxidation leads to the formation of nitrates in the same circumstances. The addition

of a dilute mineral acid to a nitrite sets free reddish-brown fumes of NO and N_2O_4. A solution of argentic nitrate forms a precipitate with cold, not too dilute, solutions of an alkaline nitrite. These two reactions distinguish these salts from the **nitrates.** The reddish fumes, above mentioned, are strong oxioxidizing agents, and set free iodine from potassium iodide. A solution of starch, with which iodine forms a deep blue color, and a solution of potassium iodide with dilute sulphuric acid, are used as a test for nitrites in solution. (Art. 122.)

Nitrous acid and the nitrites act as reducing agents upon an acid solution of potassic permanganate, and decolorize this latter salt. The nitrites can be taken up by plants, and elaborated into their structure, and hence, are valuable fertilizers.

211. **Nitrogen Peroxide, Nitrogen Tetroxide, Nitrogen Dioxide** (N_2O_4).—It may be prepared by mixing two volumes of nitric oxide with one of oxygen.

$$2NO + O_2 = N_2O_4.$$

More easily, by heating dry plumbic nitrate in a retort, passing the vapors into a cooled receiver, where they condense into a liquid:

$$Pb\,(NO_3)_2 = PbO + N_2O_4.$$

Composition.—N_2O_4 appears to exist in a pure state only at temperatures below 0°C. (32°F). The liquid is colorless at these temperatures, but at its boiling point is yellow in color, owing to partial dissociation into NO_2, which is complete at about 150°C. (302°F.)

The gas is always reddish-brown in color, due to the presence of NO_2, while N_2O_4 is colorless. The density of the gas at 26°C. (78.8°F.) (the boiling point of the liquid), is 38, and contains 20 per cent. of NO_2. On raising the temperature, the density diminishes and finally becomes constant at 150°C (302°F.), and equals 23. This density corresponds to $NO_2 = 46$. This phenomena of dissociation is frequently noticed in determining the density of bodies at temperatures much above their boiling points. The laws of quantivalence seem to hold, in these cases, only at the lower temperatures at which dissociation does not take place; hence, the confusion that exists in the formulæ of such bodies as

$$N_2O_4 \text{ or } NO_2,\ N_2O_2 \text{ or } NO,\ Hg_2Cl_2 \text{ or } HgCl, \text{ etc.}$$

212. **Properties.**—Cold water, in small quantity, decomposes it into N_2O_3 and HNO_3, while in larger quantities and with alkaline hydrates, it forms nitrous and nitric acids, or their salts.

The tetroxide, N_2O_4, and the dioxide, NO_2, both act as strong oxidizing agents, setting iodine free from the iodides.

213. **Nitric Anhydride and Acid.**—N_2O_5 and HNO_3. **Nitric Anhydride** is a white crystalline solid, fusing at 30°C. (86°F.), and boiling at 47°C. (116.6°F.). Obtained by treating dry silver nitrate with chlorine, or by the removal of water from

fuming nitric acid by the action of phosphoric oxide (P_2O_5). The oxide is unstable, has a strong affinity for water, with which it forms nitric acid. It has no especial use or interest.

214. **Nitric Acid — Aqua Fortis, or spirits of nitre** (HNO_3), is the most important of the acids of nitrogen. Does not occur free, but as nitrates widely disseminated. It is usually prepared, commercially, by the action of sulphuric acid upon potassium or sodium nitrate, in glass or cast iron retorts.

$$NaNO_3 + H_2SO_4 = HNaSO_4 + HNO_3.$$

By the application of heat to the mixture, the HNO_3 distills over and is condensed in earthenware Woulfe bottles.

It is also formed in small quantities by the passage of electric discharges through a mixture of nitrogen and oxygen. This takes place in the air by the passage of flashes of lightning, probably by the oxidizing action of the ozone generated by these phenomena. The nitrates are formed in the soil and natural waters by the oxidation of organic matter, called nitrification, and is induced by certain microscopic organisms called the nitrifying ferment. In some localities the process is conducted artificially. (See Potassium Nitrate.)

The commercial acid, prepared as above, contains sulphuric acid, traces of iron, brown nitrous oxides and chlorine. It is purified by redistillation with plumbic nitrate, which retains the impurities and allows the pure acid to distill over.

215. **Properties.**—The pure acid is a colorless, rather heavy, fuming liquid, having a sp. gr. of 1.52, boiling at 86°C. (186.8° F.), and solidifying at —40°C. (—40°F.). The sp. gr. and boiling-point of the weaker acid, vary with the proportion of acid present. When strongly heated, or on exposure to light and air, the acid is decomposed into nitric tetroxide (N_2O_4), water, and oxygen, and turns yellow.

Nitric acid readily gives up a portion of its oxygen, and thus acts as a strong oxidizing agent, attacking and destroying vegetable and animal tissues and coloring matters. It is sometimes used as a cauterizing agent, first producing a yellow stain, then destroying the tissue. While it oxidizes most organic bodies, it enters into the composition of others, forming substitution products. Thus glycerine, cotton, sugar, etc., when treated with it form explosive substitution products. Most metals dissolve in the acid, forming nitrates; gold and platinum are exceptions. The non-metals or negative elements are usually oxidized by it. Metallic iron dissolves readily in the dilute, but when plunged into strong acid, it assumes a condition known as the **passive state**; if now it be put into dilute acid, it is not attacked by it until a piece of platinum is brought in contact with it, or by some other means the passive condition is destroyed. **Nitroso-nitric** acid is a yellow, partially decomposed acid, containing nitric peroxide (N_2O_4). **Aqua regia** is prepared by mixing together four parts of hydrochloric, and one of nitric acid; it soon assumes a yellowish-red color, and has the power to dissolve gold, platinum, and other metals, forming chlorides.

Fuming Nitric Acid.—A reddish-brown acid. Sp. gr. 1.525. Containing N_2O_3 or N_2O_4. Used as a powerful oxidizing agent.

216. **Officinal Forms.—Acidum Nitricum.** Sp. gr. 1.42; contains 69.4 per cent. of HNO_3.

Acidum Nitricum Dilutum is prepared by adding six parts of distilled water to one of the above acid. Sp. gr. 1.059, and contains 10 per cent. of HNO_3. Used for internal administration. Dose ♏5–15.

Acidum Nitro-hydrochloricum, or **Nitro-muriaticum.** Prepared by mixing 4 parts of acidum nitricum with 15 parts of acidum hydrochloricum.

Acidum Nitro-hydrochloricum Dilutum; is made by adding 76 parts of water to the above formula for the strong acid (U. S. P. 1880).

217. **Tests.**—1. Add to suspected liquid some ferrous sulphate, and pour the mixture on some strong sulphuric acid in a test tube. A black, brown, or reddish zone at the point of contact of the two liquids indicates nitric acid. (See Art. 207.)

2. Heat the suspected solution with some sulphuric acid faintly colored with indigo, when, if nitric acid or a nitrate be present, the blue color will disappear.

3. The strong acid imparts a deep red color to the alkaloid, brucine.

4. When heated with copper turnings, the liquid assumes a green color, and evolves reddish fumes. When the acid is in combination, a stronger acid must be added to set it free, as in test 2.

218. **Physiological Effects.**—In small quantities, well diluted, it is a stomachic tonic, and augments the secretion of urine. It seems to be mostly decomposed in the body, but a small quantity may pass into the urine as nitrates; it acts, therefore, as an oxidizing agent. The strong acid is a corrosive, violent poison, staining the tissues and vomit with which it comes in contact

a bright yellow color, and then corroding them. These stains will be found on the tongue and fauces in cases of poisoning by this acid. **Antidote**—milk of lime, magnesia, or other alkalies well diluted, followed by sustaining treatment.

PHOSPHORUS.

$P = 31.$

Density of vapor = 62. Mol. wt. 124 = P_4.

219. **Occurrence.**—Discovered by Brandt, in 1669, in urine; and by Gahn, in bones, in 1769. Does not occur native, but as phosphates and in organic substances. Most common form is the calcium phosphate, ($Ca_3(PO_4)_2$), derived from bones of prehistoric mammals.

220. **Preparation.**—Phosphorus is usually prepared from the ash of burnt bones, in which it exists as tri-calcium phosphate, $Ca_3(PO_4)_2$. The ash is first converted into a soluble monocalcium phosphate, sometimes called superphosphate, by treating it with sulphuric acid.

$$Ca_3(PO_4)_2 + 2H_2SO_4 = CaH_4(PO_4)_2 + 2CaSO_4.$$

The $CaH_4(PO_4)_2$ is dissolved in water, and drawn off, leaving the $CaSO_4$ in the vat. This solution is evaporated to dryness, after adding powdered charcoal, and then transferred to a retort, whose beak dips under water. The retorts are then gradually heated to a high temperature, when the $CaH_4(PO_4)_2$ is first dehydrated and converted into calcic metaphosphate, $Ca(PO_3)_2$, and water, and then undergoes reduction under the action of the carbon, as follows:—

$$4Ca(PO_3)_2 + 5C_2 = P_4 + 10CO + 2Ca_2P_2O_7.$$

The free phosphorus distills over and condenses under the water as an impure article, which is purified by redistillation; or by fusing it under water with sulphuric acid and potassium dichromate, and then casting it into sticks, in moulds.

221. **Properties—Physical.**—Phosphorus is met with in several distinct allotropic states.

The ordinary form is a translucent, waxy-looking solid, which, at ordinary temperatures, is tenacious, and about the consistency of wax; but at 0°C. (32°F.) and below, it becomes brittle. It melts at 44°C. (111°F.) under water, and boils at 290°C. (554°F.) By the action of light, it soon becomes coated with a whitish or reddish coat, probably an oxide. The sp. gr. is 1.83 at 10°C. (50°F.) It shines in the dark, and when exposed to moist air, emits the odor of ozone. It is insoluble in water and alcohol, but soluble in ether, benzine, petroleum, and in the fixed and essential oils. The best solvent is carbon disulphide. From this solution, it separates in the form of octahedral and dodecahedral crystals. When a portion of the solution is poured upon filter paper and allowed to evaporate spontaneously, it takes fire when the evaporation is complete.

Red or Amorphous Phosphorus is a reddish-brown amorphous powder, of sp. gr. 2.14, insoluble in carbon disulphide; does not alter in the air, and does not show the phosphorescence in the dark. While ordinary phosphorus is very poisonous, even to workmen handling it, this variety is entirely harmless. When heated to 260°C. (500°F.), it does not melt, but gradually sublimes. The vapor is converted into the ordinary form, which takes fire in presence of air.

The red phosphorus is prepared by heating the ordinary variety for about thirty-six hours, to a temperature of from 250°C. (482°F.) to 300°C. (572°F.) in an atmosphere of hydrogen, carbon dioxide, or in an exhausted iron vessel. The mass is then washed with carbon disulphide, to remove any of the ordinary variety remaining.

Other varieties of phosphorus have been formed. The **metallic** form by heating the red variety, in a sealed tube, to 530° C. (986° F.) when black, metallic-looking, microscopic needles sublime into the cooler portions of the tube. The sp. gr. of this variety is 2.34, and it is less active than the red variety.

White Phosphorus is a form described by Remsen and Keiser as a snow-white, plastic mass, easily oxidized, formed by carefully distilling phosphorus in an atmosphere of hydrogen, and passing the vapor into a receiver containing ice water. Other forms of **white phosphorus** and **black phosphorus** have been described, but little is known of them.

Chemical Properties.—The most characteristic property of phosphorus is its ready oxidation. If the ordinary or the white variety be heated to 60°C. (140°F.) in contact with air, it takes fire and burns with a brilliant flame, and evolves a voluminous white cloud of phosphoric pentoxide. It may be burned under warm water, by throwing a jet of oxygen upon it.

It must be kept under water to prevent its taking fire spontaneously. When fragments, partly covered with water, are exposed to the air, white fumes are seen to arise from them, which contain ozone, hydric peroxide, and possibly ammonium nitrite H_4NO_2). This ozone is the cause of the odor usually detected when phosphorus is exposed to the air. The red variety does not oxidize in the air, and may be handled with impunity. Phosphorus unites readily with fluorine, chlorine, bromine and iodine, forming, in each case, two compounds of the general formula PR_3 and PR_5, except in the case of iodine, which forms PI_2 and PI_3. An oxychloride of phosphorus ($POCl$) is also known. It combines with most other elements except carbon, nitrogen and hydrogen. It reduces some metallic salts, as copper and silver, to the metallic state.

222. **Tests.**—Its phosphorescence in the dark, either as found, or after separating it with carbon disulphide and evaporation of the latter liquid. It imparts a green color to the hydrogen flame, when this gas is conducted through a solution containing it before being burned.

223. **Physiological Action.**—Owing to the ready inflammability of the ordinary variety, deep burns are liable to occur from careless handling, which are more serious and difficult to heal than burns from other combustibles. When taken internally, phosphorus is a very poisonous substance. Cases of poisoning from "ratsbane" or "rat poison" containing it, are not infrequent.

The symptoms of acute poisoning are a garlicky odor and taste in the mouth, heat and burning in the stomach, vomiting of a dark-colored matter, which is phosphorescent, when shaken, in the dark. Weak pulse, low temperature, cold extremities, dilated pupils, and a clear mind are usually seen. Death in from 2 to 12 days. Average about 3 to 4 days.

These symptoms may make their appearance in an hour after the poison is taken, or after 3 or 4 days.

Poisonous dose varies. Gr. $\frac{1}{50}$ to $\frac{1}{8}$ has been fatal.

Antidotes. There is no chemical antidote. Emetics, or the stomach pump are the best early treatment; then, mucilaginous drinks, with lime or magnesia, or oil of turpentine. The old oil is best; but no other oils should be given, as they dissolve the phosphorus and favor absorption. Recovery is rare.

Chronic poisoning of workmen in match factories frequently occurs; the symptoms are fatigue, pains in stomach and bowels, with diarrhœa, carious teeth, swollen and inflamed gums, and finally necrosis of the jaws, usually the lower. Fatty degeneration of the liver, kidneys, heart and other muscles, and destruction of the red corpuscles, are also noticed.

These evils are now remedied by using **red phosphorus** in making matches, as it is not poisonous.

224. **Phosphorus and Hydrogen — Phosphoretted Hydrogen. Phosphine.**—There are three hydrides of phosphorus known, which are all formed together by boiling phosphorus with strong potash or soda lye, or with milk of lime. They appear as a gaseous mixture, which takes fire spontaneously on coming to the air. When the beak of the retort in which it is prepared dips under water, a precaution always to be taken, each bubble ignites on coming to the surface, producing beautiful white rings of P_2O_5. This inflammable gas, composed mostly of PH_3, is found, on examination, to contain also a liquid compound, (P_2H_4), which is highly inflammable on exposure to air, while the gas (PH_3) is not. This yellow volatile liquid, on standing in sunlight, deposits a yellow solid (P_4H_2).

Phosphine (PH_3) is colorless, sparingly soluble in water, and has a strong alliaceous odor.

The impure gas is formed during the putrefactive decomposition of organic substances containing phosphorus, especially under water, and takes fire spontaneously on rising to the surface, producing the **ignis fatuus** or " Will o' the wisp " sometimes seen in marshy places. The gas is very poisonous, even in small quantities. The blood after deaths caused by it, is found to be dark colored, with a violet tinge, and has lost the power of absorbing oxygen. It poisons, therefore, by its reducing action on the blood. Its density is 17, and its sp. gr. 1.134.

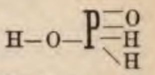
Phosphine resembles the corresponding compound of nitrogen, (NH_3), in some respects. It unites directly with HBr and HI, to form **phosphonium bromide,** (PH_4Br), **and iodide,** (PH_4I), corresponding with the ammonium compounds, NH_4Br and NH_4I.

225. **Phosphorus and the Halogens.**—Phosphorus forms three compounds with chlorine. **Phosphorus trichloride,** PCl_3, is a colorless fuming liquid, boiling at 70°C. (165.2°F.) sp. gr. 1.61, and prepared by direct union of the elements. It is much used as a reagent in organic chemistry.

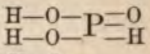
Phosphorus Pentachloride, $(PCl)_5$, is a yellowish-white, crystalline solid, fuming in the air, and subliming without fusion when heated. Prepared by treating PCl_3 with excess of chlorine, and used as a reagent in organic chemistry.

Phosphorus Oxychloride, ($POCl_3$), is formed by the action of a limited quantity of water on the pentachloride. It is a colorless liquid with a pungent odor, and a sp. gr. of 1.7. Boils at 110°C. (230°F.) Phosphorus unites directly with bromine, giving a tribromide and pentabromide, and with iodine, giving two crystalline solid compounds, PI_3 and P_2I_4, and with fluorine to form PF_3 and PF_5. These latter compounds, as well as those of phosphorus with sulphur, of which there are six known, are of no interest to the medical student.

226. **Phosphorus and Oxygen.**—The following oxides and acids are known :—

Hypophosphorous Acid .....

$\mathrm{H{-}O{-}P}\begin{matrix}=O\\ -H\\ \diagdown H\end{matrix}$

$P_2O_3 = O{=}P{-}O{-}P{=}O$
Phosphorus Oxide

Phosphorous Acid.....

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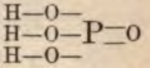
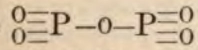
Common Phosphoric Acid......

$\left\{\begin{matrix}H{-}O{-}\\ H{-}O{-}\\ H{-}O{-}\end{matrix}\right. P{=}O$

$P_2O_5 =$

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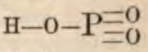
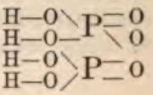
Phosphoric Oxide

Metaphosphoric Acid.

$\mathrm{H{-}O{-}P}\begin{matrix}=O\\ =O\end{matrix}$

Pyrophosphoric Acid ..

$\left\{\begin{matrix}H{-}O\diagdown\\ H{-}O{-}\\ H{-}O\diagdown\\ H{-}O\diagup\end{matrix}\right. \begin{matrix}P\begin{matrix}=O\\ \diagdown O\end{matrix}\\ P{=}O\end{matrix}$

The **sodium salt** of two other unimportant acids, the **hexabasic** and **dodecabasic acids,** are also known.

Phosphorous Oxide, or **Phosphorus Trioxide,** a white powder formed by the slow oxidation of phosphorus in dry air, is unimportant in medical chemistry. It forms no corresponding acid.

227. **Phosphoric Oxide—Phosphorus Pentoxide** (P_2O_5), is formed by the rapid burning of phosphorus in oxygen or air, and rises as a voluminous white cloud. It has a powerful affinity for water, producing a hissing noise when dropped into it, and forming a solution of metaphosphoric acid (HPO_3).

228. **Common Phosphoric Acid.**—This is the most important of these acids. Readily prepared by boiling phosphorus in diluted nitric acid, and evaporating the solution to a syrupy consistency; or, by decomposing phosphates with sulphuric acid. By spontaneous evaporation over sulphuric acid under a bell jar, hard, transparent, deliquescent, prismatic crystals are obtained. **Acidum Phosphoricum** (U. S. P.) is a colorless, non-fuming, strongly acid liquid, of sp. gr. 1.347. It should be free from arsenic, which is often present in the commercial acid.

Acidum Phosphoricum Dilutum (U. S. P.) is prepared by adding four parts of distilled water to one part of the strong acid, has a sp. gr. of 1.057 and contains 10 per cent. of H_3PO_4. This acid forms a series of well known salts called phosphates, of which the sodium, ammonium and calcium salts are the most important. The acid being tribasic, it is capable of forming acid, double and triple phosphates, of which the following are examples:—

ACID SALTS.	
Di-hydric-sodium Phosphate	H_2NaPO_4.
Hydro-calcium Phosphate	$HCaPO_4$.
NORMAL SALTS.	
Potassium Phosphate	K_3PO_4.
Calcium Phosphate	$Ca_3(PO_4)_2$.
DOUBLE SALTS.	
Ammonio-magnesium Phosphate	NH_4MgPO_4.
Potassio-barium-Phosphate	$KBaPO_4$.

229. **Pyrophosphoric and Metaphosphoric Acids,** ($H_4P_2O_7$ and HPO_3).—When ordinary phosphoric acid is heated to a temperature of about 213° C. (415°F.), two molecules lose one molecule of water, and then unite to form a doubly condensed molecule, which is called **pyrophosphoric acid.**

$$2H_3PO_4 = H_4P_2O_7 + H_2O.$$

If this new acid is heated to a temperature approaching redness, the so-called **glacial** or **metaphosphoric acid** is formed as a white, glassy, transparent, odorless solid, having a sour taste, and the formula HPO_3. $H_4P_2O_7 = 2HPO_3 + H_2O$.

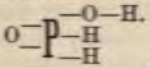
This acid corresponds in composition to nitric acid (HNO_3). It is usually prepared by heating ammonium phosphate to a red heat. It is deliquescent in the air, and very soluble in water. It has been used as a delicate test for albumin, and has the advantage of being easily handled in the solid state.

Meta- and **Pyrophosphoric Acids** and their salts, when taken internally, are said to have a decided inhibitory action upon the motor ganglia of the heart, and may even cause its cessation and death. **Pyrophosphoric** has a more decided action than the **meta- acid.**

230. **Hypophosphorous Acid** (H_3PO_2).—When ordinary phosphorus is boiled with a solution of sodium, potassium, barium, or calcium hydrate, phosphorus hydride escapes, and there is formed in solution a hypophosphite of the metal present. From the barium salt the acid may be prepared by treatment with enough dilute sulphuric acid to precipitate the barium as sulphate. The filtered solution is then to be concentrated under the air pump, as heat decomposes it. The acid, thus prepared, is a colorless, syrupy, strongly acid liquid; it is unstable in the air, gradually changing into phosphorous and phosphoric acid. The acid is of little importance, but several of its salts are used in medicine.

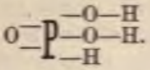
They are generally administered in the form of syrup. They have a strong reducing action on many metallic salts, and this should be remembered in prescribing them. **Mercuric chloride** is reduced to **metallic mercury** by the alkaline **hypophosphites,** and **ferric** to **ferrous salts.**

Hypophosphorous, as well as phosphorous acid, is peculiar in its composition. While there are three atoms of hydrogen in the molecules of both acids, but one in the first and two in the second are basic; *i. e.*, that can be replaced by a basic radical or metallic atom. In these two acids the phosphorus is pentad, as in phosphoric acid; and in the former the two non-basic hydrogen atoms are believed to be united directly to the phosphorus, while the basic hydrogen atom is linked to it by an oxygen atom, thus—



$$O{=}P\begin{array}{l}-O-H.\\-H\\-H\end{array}$$

In phosphorous acid we have likewise two hydroxyl groups and one non-basic hydrogen atom, thus—



$$O{=}P\begin{array}{l}-O-H\\-O-H.\\-H\end{array}$$

These acids illustrate the definition of an acid, given in Part I, viz., that in all acids, except the binary acids, basic hydrogen must be tied to the kernel element by a linking third atom, usually oxygen or sulphur; and any hydrogen not so united is not basic.

231.—**Phosphorous Acid** (H_3PO_3).—Best obtained by the action of water on the trichloride.

$$PCl_3 + 3H_2O = PO_3H_3 + 3HCl.$$

The HCl is evaporated, and the remaining solution is concentrated over sulphuric acid. It is a colorless, syrupy, highly acid and unstable liquid, and under the air-pump may be made to crystallize. It is easily decomposed by heat. It acts as a strong reducing agent by taking up oxygen, and forming phosphoric acid. It forms a series of unimportant salts, called **phosphites.**

ARSENIC.

As = 75.

232. **Occurrence.**—Arsenic occurs native and in the form of arsenides, the sulphides, orpiment and realgar, and as arsenical pyrites or mispickel. Besides occurring in these minerals in considerable quantities, it is contained in small quantities in a great number of other minerals and even organic substances. The sulphides, and even the element was known to the ancients.

233. **Preparation.**—Usually obtained in the form of the oxide by calcining mispickel, and condensing the white volatilized As_2O_3; this oxide is then strongly heated with charcoal to obtain the element, which distills over. Although this is the method most used, it may be obtained from other minerals containing it.

234. **Properties.**—A brittle, steel-gray, crystalline solid, possessing a metallic lustre, and a sp. gr. of 5.75. It also exists as an amorphous, lustreless, black mass, easily pulverized, and having a sp. gr. of 4.71. When heated out of contact with air, under ordinary pressures, it sublimes at 180°C. (356°F.) without previous fusion; but under strong pressure it fuses. Its vapor has a yellow color and a density of 150; its molecular weight is, therefore, 300, and its molecular formula As_4. In dry air it is permanent; but when heated, it burns with a bluish flame, emitting the garlic odor, and white fumes of arsenious trioxide (As_2O_3). It combines directly with many of the elements, both metallic and non-metallic, as chlorine, bromine, iodine, copper, iron, etc., yielding arsenides—the metallic arsenides resembling alloys. It combines readily with **nascent hydrogen,** which takes it from any of its compounds. Nitric and sulphuric acids are decomposed by it, without forming salts. It is oxidized by boiling solutions of caustic potash, while a portion of it is given off as hydride.

Arsenic is used in pyrotechny, in the manufacture of fly poison (under the name of cobalt), in shot, and in certain pigments.

235. **Arsenic and Hydrogen.**—Two arsenides of hydrogen are known, AsH_3, a gas, and As_2H, a solid.

Hydrogen Arsenide, Arseniuretted Hydrogen, or **Arsine,** AsH_3—is of great practical interest to the toxicologist, as it enters into some of the most delicate tests for the detection of the element. It may be prepared by a number of reactions, the most common of which are the following:—

1st. By decomposing the metallic arsenides with hydrochloric acid.

2d. By the action of hot caustic potash solutions, with metallic zinc, upon reducible arsenical compounds.

3d. By the action of nascent hydrogen upon arsenical compounds, similar to the last mentioned.

. By the reducing action of moist organic matter upon unds of arsenic.

It is a colorless gas, with a strong garlic odor, combustible in air, burning with a bluish-white flame, and emiting white fumes of As_2O_3; a cold surface pressed down upon this flame receives a black stain of arsenic. When passed through a tube heated to a dull-red heat, it is decomposed into hydrogen and arsenic, which last deposits in the cooler part of the tube as a metallic mirror. (See Marsh's Test.)

The gas is readily decomposed by oxidizing agents and the alkaline hydrates. It is exceedingly poisonous. **The solid Arsenide of Hydrogen** is of little interest.

236. **Arsenic and the Halogen Elements.**—Arsenic forms one compound with each of this group of elements, with the general formula AsR_3, in which R stands for a halogen atom.

The tri-fluoride and the tri-chloride are liquids, the first boiling at 63°C. (145.4°F.), and the second at 134°C. (273°F.). They are formed when a fluoride or chloride is heated with the trioxide and sulphuric acid.

The tri-iodide and tri-bromide of arsenic are obtained by direct union of the elements. They are both solids. The tri-iodide is used in medicine. Both these compounds are decomposed by contact with a small quantity of water; the iodide, however, dissolves in a large quantity of water. It enters into the composition of Donovan's solution, as **Liquor arsenii et hydrargyri iodidi** (U. S. P.), which is composed of one part each of mercuric iodide and arsenic tri-iodide dissolved in 100 parts of water.

237. **Arsenic and Sulphur.**—There are at least three well known compounds of arsenic and sulphur, As_2S_3, As_2S_5, AsS_2.

Arsenic Tri-sulphide (As_2S_3) occurs native, as **orpiment,** in the form of gold-yellow crystalline masses. It may be prepared by precipitating arsenious acid or its salts, with hydric sulphide, or by heating sulphur and arsenious oxide (As_2O_3). It is lemon-yellow in color, soluble in ammonium hydrate and the alkaline hydrates, and in yellow ammonium sulphide, but insoluble in water and acids.

Arsenic Pentasulphide (As_2S_5) is also a bright yellow powder of no special interest.

Arsenic Disulphide (AsS_2) occurs native as **realgar** in the form of ruby red crystalline masses of sp. gr. 3.5. Realgar and orpiment are used as pigments.

238. **Oxides and Acids of Arsenic.**—Arsenic forms two oxides, with corresponding acids:—

As_2O_3 **Arsenious Oxide, or Anhydride.**	H_3AsO_3 **Arsenious Acid.**
As_2O_5 **Arsenic Oxide, or Anhydride.**	H_3AsO_4 **Arsenic Acid.**

Arsenious Oxide (As_2O_3) is the most important of the compounds of arsenic. It occurs in nature as arsenic "bloom." It is obtained artificially as a side product in roasting ores of other metals containing arsenic, when it volatilizes, and is condensed in large chambers as a white powder. It is purified by resublimation in iron retorts, and is obtained in the form of a white powder or glassy-looking solid, of a sp. gr. of 3.69.

239. **Properties.**—As ordinarily met with it (white arsenic) is a white, somewhat gritty powder, which, under the microscope, is seen to be made up of more or less regular octahedral crystals. When the vapor is rapidly cooled, the crystals take the form of rhombic prisms, and it is, therefore, **dimorphous.** When heated, it sublimes without fusing, at about 218°C. (424°F.) When heated in sealed tubes, it melts into a vitreous mass. The density of the vapor is 198, corresponding to the formula As_4O_6, which is probably the formula of the vitreous variety, while that of the octahedral variety is As_2O_3. It is soluble, with difficulty, in water, forming a sweetish, metallic, and nauseous-tasting, poisonous solution of arsenious acid (q. v.) It is also soluble, with decomposition, in hydrochloric acid and alkaline solutions, playing the basic rôle in the former, and the acid rôle in the latter. Nascent hydrogen reduces the oxide, and converts the arsenic into arsine (AsH_3), while oxidizing agents convert it into arsenic acid.

Antidote.—Freshly prepared ferric or magnesium hydrate, or a solution of **dialyzed iron.**

240. **Arsenious Acid,** (H_3AsO_3), is formed by dissolving As_2O_3 in water. A solution of the acid in dilute hydrochloric acid is officinal under the name of **liquor acidi arseniosi** (U. S. P.), containing 1 per cent. of As_2O_3. It forms a series of salts called arsenites. **Potassium arsenite** is officinal as Fowler's solution, or **Liq. Potass. Arsenitis.**

Scheele's green is an arsenite of copper, used as a pigment.
Paris green is a mixture of acetate and arsenite of copper.

241. **Arsenic Oxide,** (As_2O_5), is a white, amorphous, deliquescent solid, dissolving in water to produce **arsenic acid.**

$$As_2O_5 + 3H_2O = 2H_3AsO_4.$$

Arsenic acid is usually prepared by warming arsenious acid with nitric acid, when the As_2O_3 is oxidized at the expense of the nitric acid. On evaporating, the solution yields needle-shaped crystals of H_3AsO_3. The aqueous solution is strongly acid. On heating the crystals of arsenic acid, both the pyro-arsenic and meta-arsenic acids are produced, corresponding to the similar acids of phosphorus. But one salt of this acid is officinal, the **Liquor Sodæ Arseniatis.**

242. **Arsenic Poisoning—Toxicology.**—From the earliest history of arsenic, it has been used as a poison for criminal purposes.

While every physician should not undertake the analysis in cases of suspected poisoning by this agent, a knowledge of the outlines of the chemist's methods of analysis, etc., will teach the physician to prepare the way for the analyst. Moreover, a few preliminary tests by the physician may frequently save much unnecessary litigation and expense in some suspected cases. Other reasons might be given why every physician should have some knowledge of toxicological science, that care may be taken to punish the guilty and protect the innocent. The attending physician is responsible for the connection of the links of evidence. All compounds of arsenic are poisonous, and the poison usually enters the system by the mouth, although it has been absorbed by the skin, mucous membrane, or abraded surfaces, in sufficient quantities to produce poisonous results, especially chronic poisoning.

Colored wall paper, colored toys, confectionery, and certain aniline dyes used in fabrics, may give rise to accidental poisoning.

243. **The Physician's duty in cases of Poisoning** may be briefly stated as follows: In case foul play is suspected, do not fail to make careful notes, at the time, as to dates, symptoms and circumstances, or facts leading to such suspicion.

The physician should collect and preserve the urine, fæces, vomit, and the suspected vehicle of the poison, and place them under seal, or lock and key. He should test some one or all of these, to satisfy himself as to the truth or falsity of his suspicion. As little publicity as possible should be given to matters of fact or opinion at the time. Be not too ready to express your opinion upon the origin of the poison in cases of this kind, lest you jeopardize the reputation of your patient or others. Whether a fatal termination is expected or not, it is wise to take these precautions.

In case of fatal termination, notify the prosecuting officer or

[illegible]

[illegible] must be the original, not a copy [illegible] at the time at which they refer or immediately after.

[illegible] *Symptoms of Arsenical Poisoning.*—The [illegible] are those of an irritant poison. There is usually marked [illegible] increased by pressure. [illegible] burning pains in the limbs and painful cramps [illegible] are usually present. Intense thirst, dry [illegible] small, rapid pulse, anxious, pinched countenance [illegible] and staring, tongue dry and furred, prostration, great restlessness [illegible] with a perfectly clear mind are symptoms usually to be expected. The urine is diminished, with frequent and painful micturition. These symptoms may end in convulsions, tetanus, collapse, or coma and death. Minimum fatal dose from 1.5 to 2.5 grains. Chronic arsenical poisoning is usually attended by conjunctival inflammation, irritation of the skin with a vesicular or nettle-rash eruption, similar to that of scarlet fever. Irritation of the stomach and bowels, exfoliation of the cuticle of skin and tongue, and falling of the hair, have been noticed. Local paralyses, preceded by numbness, or tingling of the toes and fingers, and marked disorders, are of common occurrence.

245. **Treatment.**—Remove any unabsorbed poison from the stomach by emetics or the stomach pump.

The best **antidote** is freshly precipitated ferric hydrate, prepared by adding aqua ammoniæ, in slight excess, to a solution of ferric sulphate or ferric chloride. The solution of **dialyzed iron,** now found in the shops, may be used instead of the above, and may be given in teaspoonful doses at short intervals. This forms an insoluble compound with the arsenious acid, and thus prevents further absorption.

The symptoms caused by the absorbed poison are to be treated as they arise.

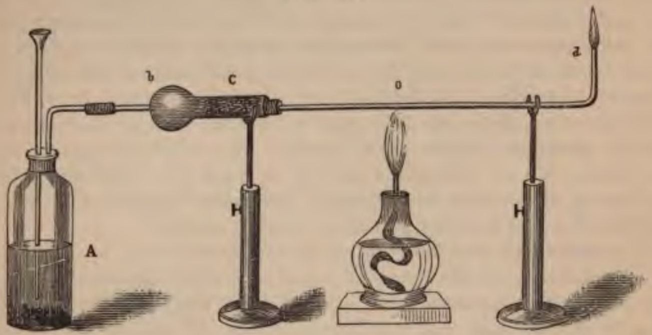
246. **Tests.**—There have been devised a large number of tests for the detection and identification of arsenic. Some of these are easy of application, while others will be used only by the chemist. We can only give the outline of these tests, and leave the student to consult special works for minutiæ. The analysis, after death, in cases of suspected poisoning, should not be undertaken by the physician, nor even by a chemist, unless he has a well equipped laboratory in which he can conduct the analysis, from beginning to end, alone, and without interruption. During the life of the patient, the physician should be able to test the urine, fæces, or suspected articles of diet, medicines, etc., for the presence or absence of arsenic. For this purpose he may use Reinsch's or Marsh's test, but it must be understood that neither of these tests alone, when performed as about to be described, is to be relied upon as positively certain.

Reinsch's test may be conducted as follows:—

To a portion of the urine or other suspected liquid add about one-sixth its volume of pure hydrochloric acid and some strips of pure copper foil, and boil the solution. If arsenic be present, a steel-gray or bluish deposit will be formed on the surface of the copper. This deposit is not positive proof, however, as antimony, bismuth and mercury may give similar deposits; it is, therefore, necessary to apply tests to this deposit, in order to determine its identity. For this purpose the copper is removed, washed in water, dried between folds of filter paper, and placed in a clean, dry, wide test tube, and heated to a dull red heat, taking care to heat only that portion of the tube where the copper is. The tube may then be broken and the inner side of the fragments examined with a microscope, for octahedral crystals of arsenious oxide. The copper and the acid used must be shown to give no stain on prolonged boiling with distilled water. This blank experiment must always be performed.

Marsh's test may be conducted as follows:—Although the indications afforded by it are not conclusive in the presence of organic matters, it should always be used to confirm Reinsch's test. Into a flask holding about 150 c.c. (f℥ v), introduce some pieces of zinc, free from arsenic and antimony; then pour over these some water acidulated with sulphuric acid; close the flask with a cork containing a funnel tube, and a delivery tube drawn to a fine point and containing a pledget of cotton in the end at the cork, arranged as in Fig. 35. After allowing the generation of hydrogen to go on for a considerable time, to expel the air from the upper part of the flask, say half an hour, light the gas at the open end of the delivery tube, and press a cold porcelain surface down upon the flame. If the materials used are free from arsenic and antimony, there will be no black stain produced on the porcelain.

FIG. 35.



Having determined that the apparatus and materials are free from impurities, put out the flame and pour the suspected fluid through the funnel tube, so as to admit little or no air with it into the flask. Now ignite the gas again, and test the flame again with the cold porcelain surface. A brilliant black or brown stain, soluble in a solution of sodium hypochlorite, is probably arsenic. Moisten one of these spots with nitric acid, when it should disappear; evaporate the acid over a lamp, moisten the spot with water, and hold the dish over a vessel containing sulphuretted hydrogen, prepared by the action of sulphuric or hydrochloric acid upon sodium or potassium sulphide. If the stain was due to arsenic the spot will turn lemon-yellow. The anti-
ny mirror is insoluble in sodium hypochlorite (Labarraque's

solution) and after treatment as above gives an orange stain. Now soften the glass, bend delivery tube downward and let it dip into a solution of $AgNO_3$; after an hour pour some NH_4OH upon the surface of $AgNO_3$ solution. Yellow precipitate at line of separation = arsenic. If the substance to be tested is a solid, a small portion of it may be thrown upon a glowing charcoal, when arsenic, if present, will give an odor resembling garlic. These tests will be sufficient to enable the physician to decide upon the presence of arsenic during the life of the patient, and guide him in his treatment and behavior.

247. **Toxicological Analysis.**—The chemist who undertakes the analysis in cases of supposed poisoning has no easy task. He must not confine his tests to any one poison. Poisons are generally divided into two groups, **inorganic** and **organic.** The limits of this work will not allow us to describe the details of procedure for the chemist, who will consult special works on toxicology.

In searching for mineral poisons, the organic matters must first be destroyed or separated by dialysis. For the latter process see Part I, Art. 72.

For the destruction of organic matter, two methods may be used. The solid matter—the stomach, and other organs—are to be cut in fine pieces and placed in a new porcelain dish, mixed with hydrochloric acid, and heated over a water-bath. Small quantities of pure potassium chlorate are put in from time to time and stirred, until the organic matter is destroyed. Or, sulphuric and nitric acids are used to thoroughly char the organic matter, and the whole diluted with water and filtered. The metals, except lead and barium, pass into the filtrate, and can be detected by either of the methods given above or by other methods. These solutions may be treated with hydrosulphuric acid gas from 24 to 48 hours, when copper, lead, bismuth and mercury give black or brown precipitates; cadmium, sometimes tin, arsenic give a yellow, and antimony an orange precipitate. In searching for arsenic, the yellow precipitate obtained, is separated from the liquid by filtration. A portion of this is to be preserved in a sealed glass tube. The remainder is oxidized with nitric acid, fused with sodium carbonate and nitrate, and the sodium arseniate thus formed may be made to yield silver and copper arseniates; the former a reddish-brown, and the latter a green precipitate. Other portions are converted into the octahedral crystals of trioxide, the metallic state, etc., the object being to present the poison in court in as many different states as possible, so as to avoid the possibility of doubt in the minds of the jury. If the yellow sulphide, soluble in ammonia, the black metal (so called), the octahedral crystals, the mirrors with the above mentioned properties, the coated copper obtained by Reinsch's test, the black deposit from silver nitrate in Marsh's test, with arseniate in the filtrate, the arsenite and arseniates of silver, and arsenite of copper, etc., etc., are obtained, with proper precautions, there is no room for doubt that the substance is really arsenic.

For further directions on this subject, the student is referred to Taylor on poisons, or Woodman and Tidy on Forensic Medicine and Toxicology, or some other similar work.

ANTIMONY.

Stibium, Sb = 122, Sp. gr. 6.71. Melting point 450°C. (842°F.)

248. **Occurrence and Preparation.**—Antimony occurs native; but the principal source is the trisulphide, Sb_2S_3, called stibnite. It occurs in a number of other ores as sulphide, or oxide.

The element is easily obtained by roasting the sulphide, and then fusing the oxide thus obtained, with charcoal.

249. **Properties.**—Antimony is a bluish-white, brittle, crystalline solid, isomorphous with arsenic and red phosphorus, and resembling zinc in color and lustre. Tarnishes with difficulty, but takes fire at a red heat. It unites readily with chlorine, forming two chlorides, $SbCl_3$ and $SbCl_5$, both of which are decomposed by excess of water. In physical properties it resembles metals, and forms alloys. In chemical properties it plays both the positive and negative rôles with facility. It is used as a constituent of type metal, Babbitt's anti-friction metal, Britannia, etc., to give hardness, and to cause them to expand and completely fill the moulds in cooling.

250. **Hydrogen Antimonide, Stibine, Antimoniuretted hydrogen** (SbH_3).—A colorless, odorless gas, formed in the same conditions as the corresponding compound of arsenic; *i. e.*, by nascent hydrogen on reducible antimony compounds. It differs from that compound in being much less poisonous, and giving a different reaction with solutions of silver nitrate, as shown in the following reactions:—

$$6\,AgNO_3 + AsH_3 + 3H_2O = 6\,HNO_3 + H_3AsO_3 + \underset{\text{Metallic Silver.}}{3Ag_2}$$

$$3\,AgNO_3 + SbH_3 = 3HNO_3 + \underset{\text{Silver Antimonide.}}{SbAg_3}$$

By carefully floating a solution of ammonium hydrate over the first solution, after passing the gas through it for some time, a yellow precipitate of arsenite of silver will be formed at the line of separation of the two liquids, while, in the case of antimony, no such precipitate will be formed. This gas is formed in Marsh's apparatus, and it is likely to be confounded with arsenic unless very great pains be taken to avoid it.

The mirrors obtained on the porcelain or in the delivery tube, require a higher temperature to volatilize them in the case of antimony; they are difficultly soluble in hypochlorite solution, are sooty and less brilliant in appearance, there is no garlic odor, and by oxidation they do not form crystals.

251. **Chlorides, Bromides and Iodides.**—Two chlorides and two oxychlorides exist.

Antimony Chloride, or **A. Trichloride—A. Protochloride—Butter of Antimony** ($SbCl_3$), may be obtained by dissolving the trisulphide in hydrochloric acid. At low temperatures it is a crystalline solid, and melts at 73.2° C. (164° F.) to a yellow, oily liquid. A solution of sp. gr. 1.47 is sometimes used as an escharotic. On the addition of considerable water, this chloride is decomposed into the **oxychloride** ($SbOCl$), formerly called **powder of algaroth.** $SbCl_3$ is poisonous, acting both locally and as a true poison.

Antimony Pentachloride ($SbCl_5$) is a fuming, colorless liquid, of little interest to the medical student. The iodides, bromides and fluorides, are similar to the trichloride in composition. The iodide has been used in medicine.

252. **Sulphides of Antimony.**—Two sulphides are known and several oxysulphides. **Antimony Trisulphide—Sulphuret of Antimony—Black Antimony—Antimonii Sulphidum** (U. S. P.), Sb_2S_3, occurs native as a steel-gray, crystalline solid. Artificially, it may be prepared by precipitating a soluble antimony salt with hydrogen sulphide as an orange-colored powder. When the native ore is roasted in air it is partially decomposed, and fuses into a vitreous, somewhat transparent mass, known as **glass of antimony,** or **crocus.**

Antimonii Sulphuretum (U. S. P.) is a reddish-brown powder prepared by dissolving the native sulphide in a solution of sodium hydrate, and re-precipitating the hot solution with sulphuric acid. It contains a little antimonious oxide (Sb_2O_3), and is used mostly in pill form. Dose gr. ij to xx. Plummer's Pills, **Pil. antimonii comp.** contain this sulphide with calomel.

By treating a hot solution of the trisulphide with sodium hyposulphite, a fine red precipitate of oxysulphide is obtained, which is used as a pigment, under the name of **antimony vermilion.** **Kermes' mineral** is another oxysulphide.

Antimonium Sulphide (Sb_2S_5) is best obtained by decomposing sulphantimoniates with a dilute acid. It is an orange-red or brown powder, readily soluble in alkalies and alkaline sulphides, forming antimoniates with the general formula $M_3'SbO_4$, in which M stands for a metallic atom. Sulphantimoniates of silver, lead, and iron, occur as minerals.

253. **Oxides and Acids of Antimony.**—Three oxides are known.

Antimony trioxide...... Sb_2O_3, antimonious acid $HSbO_2$.
" **pentoxide...** Sb_2O_5.
" **tetroxide....** Sb_2O_4, or $SbO.SbO_3$.

Antimony Trioxide (Sb_2O_3) is obtained by roasting the metal in air, or by treating it with HNO_3 and evaporating excess of acid. It is dimorphous and crystallizes in the same forms as As_2O_3, and is therefore isomorphous with it. It is a white powder capable of being sublimed.

Antimonious Acid or **Hydrate** ($HSbO_2$) is obtained by adding a solution of sodium carbonate to a solution of $SbCl_3$, as a white precipitate.

$$2\,SbCl_3 + 3\,Na_2CO_3 + H_2O = 2\,SbO_2H + 6\,NaCl + 3\,CO_2.$$

By boiling, this hydrate is changed into the trioxide. It reacts with both acids and alkalies to form salts. Thus, we have $NaSbO_2$, $SbO(NO_3)$ and $Sb(NO_3)_3$. We also have **antimony sulphate** ($Sb_2(SO_4)_3$—and **antimonyl sulphate** $(SbO)_2(SO_4)$—the former by dissolving the oxide in strong, and the latter in dilute sulphuric acid. Both are decomposed by excess of water.

Antimonic Acid ($HSbO_3$)—is obtained, by treating antimony with warm concentrated HNO_3, as a white powder, insoluble in water and nitric acid.

Pyro-antimonic Acid ($H_4Sb_2O_7$)—is also known, and may be obtained by treating its salts with hydrochloric acid.

By gently heating either of the above acids, antimony pentoxide is obtained as a yellow, amorphous mass, and by a stronger heat, with free access of air, it is converted into the tetroxide—Sb_2O_4—which is usually regarded as an antimoniate of antimonyl ($SbO\ SbO_3$). It is a white, non-volatile powder, becoming yellow when heated.

254. **Potassium Antimonyl Tartrate.—Tartar Emetic—Antimonii et Potassii Tartras** ($KSbOC_4H_4O_6.H_2O$) is one of the most commonly employed compounds of antimony.

Prepared by boiling 3 parts of Sb_2O_3 with 4 parts of cream of tartar ($KH\,C_4H_4O_6$) in water, for an hour, filtering, evaporating the filtrate and allowing it to crystallize out. It occurs in small, transparent, rhombic crystals, which effloresce in air, and have a sweetish, afterward disagreeable, metallic taste and acid reaction. Soluble in 17 parts of water at 15°C. (59°F.), and 3 parts boiling water; insoluble in alcohol. Its solutions are incompatible with alcohol, hydrochloric acid, and alkaline carbonates. Tartaric acid prevents the precipitates caused by the above reagents. On being heated to redness it chars.

It is used in medicine, and enters into the composition of **Syr. scillæ compositus, Vinum antimonii** and **Unguent. antimonii.** The dose of tartar emetic is gr. j–ij (.065–.125 grms.), as an emetic; as an expectorant, gr. $\frac{1}{16}$ to gr. $\frac{1}{4}$ (.004–.016 grms.); of the wine, 10 to 30 drops.

255. **Physiological Action.**—Locally, the soluble compounds of antimony act as powerful irritants. Tartar emetic causes a pustular eruption resembling variola, which is accompanied with fever and systemic results. Cases of poisoning from antimony used in dyeing clothing have been reported. Internally, tartar emetic is employed as an expectorant, sudorific, sedative, nauseant and emetic, according to the dose used. In full doses it causes vomiting, purging, griping pains, with great depression. In excessive quantity it acts as an irritant poison, and has produced death by syncope, preceded by convulsions and delirium. One and a half grains (.092 grms.) have produced death, but recovery has occurred after very large doses, because of the rejection of the poison by vomiting.

The treatment should consist in promoting free vomiting, or removal of the poison with the stomach pump. The proper antidote is tannic acid, which forms an insoluble compound with antimony; it may be administered in the form of infusion of tea, oak bark, nutgalls, etc., which contain it. Stimulants are then to be administered.

In suspected cases, examine the urine or viscera by Marsh's test (Art. 246). Soluble salts of antimony give an orange-colored precipitate with H_2S, in acid solutions, which is soluble in yellow ammonium sulphide, and in strong hydrochloric acid (HCl). In Reinsch's test, a bluish stain is obtained, but the sublimate obtained from it is amorphous, not crystalline. (See Art. 246.)

BISMUTH.

Bi = 210.

256. **Occurrence and Preparation.**—Occurs native and as sulphide—bismuthinite. The element is obtained by roasting the sulphide in air, and reducing the resulting oxide with charcoal.

257. **Properties.**—Bismuth is a white, metallic-looking solid, with a bronze tint. Sp. gr. 9.9. Brittle, and crystallizes in rhombohedrons; fuses at 267°C. (512.5°F.), volatilizes at a white heat, and if heated in air it burns to Bi_2O_3. HNO_3 and hot H_2SO_4 dissolve it, but HCl does not. Water precipitates basic salts from the solutions of the neutral salts. It alloys with the metals, and is sometimes described as a metal.

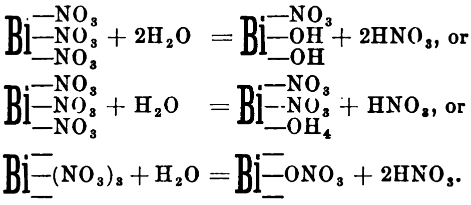
258. **Bismuth Chloride** ($BiCl_3$) may be obtained by treating the element with chlorine or aqua regia. It is a soft, white, deliquescent, volatile solid. Water added to its solutions precipitates the white **oxychloride.**

$$BiCl_3 + H_2O = BiOCl + 2HCl.$$

This reaction resembles that of $SbCl_3$. BiOCl, as well as the subnitrate, is sold as pearl powder, or pearl white, and used as a cosmetic. They blacken by H_2S. The compounds $BiBr_3$ and BiI_3 are similar to $BiCl_3$. Bismuth does not form a hydride (BiH_3), as do the rest of the group.

259. **Oxygen Compounds.**—Bismuth oxide (Bi_2O_3) is a yellow powder, insoluble in water and alkalies; it may be prepared by roasting bismuth, or heating the nitrate or carbonate. When chlorine is passed through a solution of potassium hydrate, in which Bi_2O_3 is suspended, **bismuthic acid** (BiO_3H or $H_4Bi_2O_7$) is precipitated as a red powder. On gently heating this the pentoxide, or **bismuthic oxide** (Bi_2O_5), is formed. Bismuth hydrate (BiO_3H_3) is not known, but a metahydrate (BiO_2H) is precipitated when caustic soda or **potassa** is added to a bismuth solution, or to the nitrate suspended in water. This oxide is the one first precipitated in testing for sugar in Bœttger's test.

260. **Bismuth Nitrate** ($Bi(NO_3)_3$) is formed by dissolving bismuth or the basic nitrate in nitric acid and evaporating, when it crystallizes in large transparent tables ($Bi(NO_3)_3 5H_2O$). It is soluble in a little water, but is decomposed by a large amount into a basic or subnitrate.

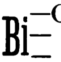


$$\mathrm{Bi}\begin{array}{l}-NO_3\\-NO_3\\-NO_3\end{array} + 2H_2O = \mathrm{Bi}\begin{array}{l}-NO_3\\-OH\\-OH\end{array} + 2HNO_3, \text{ or}$$

$$\mathrm{Bi}\begin{array}{l}-NO_3\\-NO_3\\-NO_3\end{array} + H_2O = \mathrm{Bi}\begin{array}{l}-NO_3\\-NO_3\\-OH_4\end{array} + HNO_3, \text{ or}$$

$$\mathrm{Bi}\begin{array}{l}-\\-(NO_3)_3\\-\end{array} + H_2O = \mathrm{Bi}\begin{array}{l}-\\-ONO_3\\-\end{array} + 2HNO_3.$$

Bismuthi Subnitras. (U. S. P., Br. P.)—The subnitrate of bismuth is not a definite and fixed compound, but a mixture. It is a white powder, insoluble in water, but soluble in nitric acid. As arsenic and bismuth frequently occur together, the latter is apt to be contaminated by the former. Should unpleasant effects arise from its use, it should be tested with one of the tests described under arsenic (*q. v.*) It is used internally, and as a dressing for wounds.



261. **Bismuth Carbonate — Bismuthi Subcarbonas** (U. S. P.)—**Bismuthi Carbonas** (B.P.)— $\mathrm{Bi}\begin{array}{l}-O\\=\end{array}CO_3$, is a light, white, odorless and tasteless powder, insoluble in water, and formed by the action of alkaline carbonates upon solutions of bismuth. Heated, it is changed into Bi_2O_3 and carbon dioxide.

Bismuth Sulphate ($Bi_2(SO_4)_3$) is formed by dissolving bismuth in sulphuric acid. It is of no special interest to physicians.

262. **Bismuth Citrate** ($BiC_6H_5O_7$) is a white, amorphous powder, odorless and tasteless, insoluble in water or alcohol, but soluble in ammonium hydrate. Prepared by boiling $BiONO_3$ in a solution of citric acid, and precipitating the citrate with water. By dissolving the citrate in dilute ammonia water, and cautiously evaporating to a syrup, and spreading on glass, shining, pearly, or translucent scales of **Bismuth et Ammonii Citras** are obtained, soluble in water. A solution of 2 pts. $BiONO_3$, and 4 pts. $KNaC_4H_4O_6$, in a strong solution of sodium hydrate is used as a delicate test for diabetic sugar. It becomes black from reduction of bismuth.

263. **Physiological Action.**—The bismuth salts in medicinal doses are tonic, antispasmodic, mildly astringent and antifermentative. They are used to allay gastro-intestinal irritation and diarrhœa. When administered in considerable quantities they form black stools, from the presence of the sulphide formed by the H_2S of the intestines. In many cases of excessive diarrhœa, with acid fermentation in the stomach and intestines, this black-

[illegible]ing does not occur, and its appearance marks an improvement in the case.

Cases of poisoning by large doses of the salts of bismuth are generally, if not always, due to the presence of arsenic in them.

264. **Tests.**—Water precipitates, in the absence of free mineral acids in considerable quantities. Hydrogen and ammonium sulphides give a black precipitate, insoluble in water, dilute acids and alkaline sulphides. **Potassium, Sodium** and **Ammonium Hydrates** give a white precipitate, insoluble in excess of the reagents. Infusion of nutgalls gives an orange precipitate. **Potassium Iodide** gives a brown precipitate, soluble in excess of reagent. A piece of paper dipped in a solution of potassium sulphocyanide and dried, forms a yellow spot when a drop of a solution containing bismuth is dropped on it.

BORON.

B = 11.

265. **Occurrence** and **Preparation.**—An unimportant element, never occurring native, but as borates and boric acid. Borates of calcium, magnesium and sodium (borax) occur native; the last, the most important, is found in India and California. The element may be prepared in two allotropic states; the first, by fusing its oxide with sodium or potassium, as a greenish-brown powder; the second, by fusing the oxide, chloride, or fluoride with aluminium, as a crystalline transparent solid, varying in color from colorless to a garnet red. Boron combines directly with nitrogen at elevated temperatures.

266. **Boric Anhydride and Acids.**—B_2O_3 is a transparent, glass-like mass, obtained by heating boric acid to redness. It is used in blowpipe analysis.

Boric or **Boracic Acid** (H_3BO_3)—exists native in lagoons, in the vicinity of volcanoes in Tuscany. By evaporation and crystallization the acid is obtained. **Borate of sodium** or **borax** occurs in California and India, and from it the pure acid is prepared by precipitating it with HCl from a hot solution. The acid separates in white, shining scales; it is soluble in 25 parts of H_2O at 14° C. (57.2° F.), and in 3 parts of boiling H_2O. The solution has a faint acid reaction. Turmeric paper moistened with the solution assumes a reddish-brown color on drying. The acid is soluble in 6 parts of alcohol, and is also soluble in glycerine, to the flame of both of which it imparts a distinct green color; this and the action on turmeric paper are used as tests.

When boiled with glycerine, an ether is formed, known as **boroglyceride**, which is soluble in water, has a neutral reaction, is tasteless, and is used as a preservative for foods. Its use, as well as that of boric acid, is attended, in considerable doses, with an increased excretion of urea with irritation of the kidneys, and should be used with some caution. Owing to its antiseptic action, it is used in surgical dressings.

When heated, H_3BO_3 loses one molecule of water at 100° C. (212° F.), and forms **metaboric** acid (HBO_2); on further heating it forms **tetraboric** acid ($H_2B_4O_7$), and at a higher temperature **boric anhydride** is formed.

SILICIUM—SILICON.

Si = 28.

267. **Occurrence.**—In native rocks, either as silicic oxide (SiO_2), quartz, amethyst, carnelian, etc., or combined with various metallic oxides as silicates. Clay is principally a silicate of aluminium colored with iron and vegetable matter. This, next to oxygen, is the most abundant of the elements. Neither the element nor its compounds are of much interest to the medical student.

The element never occurs native, and may be prepared in three allotropic states: **amorphous silicon, graphitic silicon,** and **crystallized silicon,** somewhat resembling the three states of carbon.

268. **Compounds.—Silicic Hydride** (SiH_4) is obtained as a colorless, spontaneously inflammable gas, by the electrolysis of a solution of common salt, using for the positive electrode an aluminium containing silicon. Silicic chloride ($SiCl_4$) is a colorless, volatile liquid possessing an irritating odor. The bromide, $SiBr_4$, and fluoride, $SiFl_4$, are also known. This latter is decomposed by water, forming **hydro-fluo-silicic acid,** H_2F_6Si.

269. **Silicic Oxide, or Anhydride** (SiO_2), is the only known oxide of this element, and exists in a pure state in quartz crystal. Artificially, it may be prepared by adding hydrochloric acid to a concentrated solution of soluble or water glass, filtering, washing and heating the residue, to expel the water. Artificially prepared it is a fine, white, tasteless powder, fusible with great difficulty, and not sensibly soluble in water or acids, with the exception of hydrofluoric acid. Its sp. gr. is 2.66. When fused with potassium or sodium carbonates or hydrates, it forms a silicate of these metals, or glass; when these alkalies are in excess, the glass is soluble in water, the degree of solubility increasing with

the proportion of alkaline salt used. This compound is known as **soluble** or **water glass.**

270. **Silicic Acid and Silicates.**—The normal silicic acid has the formula $H_4S_2O_4$, and is only known in solution in water. It may be prepared by adding hydrochloric acid to a very dilute solution of an alkaline silicate, but it is unstable. The acid is very prone to liberate a portion of its water and form acids of the condensed types; *i. e.*, two or more molecules unite and liberate one or more molecules of water. The native silicates are very complex in structure, and are usually formed on this condensed plan.

271. **Glass.**—Common glass is a mixture of several silicates, in which there is an excess of silica, the principal ones used being sodium, calcium and lead silicates. By the addition of small quantities of metallic oxides, various colors are imparted to the glass, as desired; thus, cobalt gives a blue, manganese an amethyst, copper a ruby or bluish-green, chromium a greenish-yellow, ferric oxide a brownish-yellow or black, ferrous oxide the ordinary bottle-green, etc.

CARBON.

$C = 12.$

272. **Occurrence.**—Carbon occurs native in the diamond, graphite (black lead), and the various forms of coal. In combination it occurs in all the organic bodies, petroleum, fats, oils, and in native carbonates (marble, dolomite, etc.).

273. **Varieties.**—Carbon exists in three allotropic states, the diamond, the graphite, and amorphous carbon. The **diamond** occurs in alluvial deposits in Brazil, India, Borneo, South Africa, and in small quantities in other localities. The so-called California diamonds and Brazilian pebbles are crystals of quartz (SiO_2). The diamond is pure crystalline carbon, possessing a brilliant lustre and a high power of refraction, and is the hardest substance known. It crystallizes in rhombohedra of the first system, but cleaves readily into octahedra. It occurs colorless, as well as colored through all shades of yellow, brown and black. The sp. gr. is 3.5. Heated in the oxyhydrogen blowpipe flame, or in oxygen gas, it burns to carbon dioxide, and is slightly softened by the heat of the electric arc.

Graphite, or **Plumbago,** occurs as a native mineral, having a grayish-black color, a lustre almost metallic, and a soapy feel, and leaves a black streak upon a white surface, on account of which it is used to make lead pencils, and is called **black lead.** It sometimes occurs in short, six-sided prisms. It burns in an atmosphere of oxygen with more difficulty than the diamond, furnishing carbon dioxide, and leaving from 2 to 5 per cent. of ash. Some of the purer varieties can be made directly into pencils, but this is not common. The poorer varieties are ground to a powder, heated with potassium chlorate ($KClO_3$) and sulphuric acid (H_2SO_4), or with strong nitric acid washed with water, heated to red heat, and then mixed with some adhesive material and pressed into cakes. These cakes are then sawed into suitable strips and mounted.

Graphite conducts heat and electricity well. It is also used in the manufacture of crucibles, as a lubricant for very heavy machinery, and as a stove polish. It may be obtained artificially by fusing amorphous carbon with cast-iron; on cooling the mixture and dissolving the iron in dilute acid (HCl), the graphite is left in the form of minute hexagonal plates. **Amorphous carbon** is found native in the form of coals of many varieties. Artificially, it is prepared by the partial burning, or carbonizing, of organic matter, such as wood, pitch, blood, etc.

Soot, or **Lampblack**, one of the purest forms of amorphous carbon, is prepared by the imperfect combustion of turpentine, pitch, or heavy oils, rich in carbon. It is used principally as a pigment. **Animal charcoal** or **bone black** is obtained by the carbonization of animal matter from slaughter houses (blood, bones, etc.), and possesses in a remarkable degree the power of removing the coloring matter or putridity from organic solutions filtered through it. It is used in the refining of sugar to remove the color from the solution of raw sugar. **Carbo animalis purificatus** is prepared by treating the above with dilute HCl, to remove the calcium phosphate. **Charcoal, carbo ligni** (U. S. P.), is carbonized wood, retaining, usually, the form and grain of the wood and the ash contained in the same. In preparing it, the wood, cut into suitable lengths, is piled up on end into a conical heap, a trough made of boards being laid from the circumference on one side to the centre, to supply air. The whole heap is then covered, first with straw or leaves, and finally with dirt about six or eight inches in thickness, except a hole at the centre on the top. The fire is lighted and is allowed to burn slowly, until the flame almost ceases to come from the hole at the top, or until the gaseous products have all burned off. This opening at the top, as well as the draught holes, is now covered over and the fire allowed to smoulder for a time, when the heap is torn down and the remaining fire extinguished with water.

It is used as a fuel, in the manufacture of gunpowder, and sometimes in the construction of filters, as it possesses, in a feebler degree than animal charcoal, the power of destroying noxious odors, and filtering coloring matters from organic solutions. 1 volume of it absorbs 90 volumes of NH_3, 55 volumes of H_2S and 9 volumes of oxygen at 100° C. (212° F.). It is very porous, and burns with little flame. **Coke** is the porous mass left in the retorts from the destructive distillation of mineral coal in the production of illuminating coal gas. **Gas retort carbon** is a compact, hard mass found adhering to the inside of the retorts in the above process of manufacturing coal gas. It has a metallic lustre at times, and is a good conductor of electricity, and is used for the negative plate in the construction of many forms of the galvanic battery.

Mineral coals (anthracite, bituminous, brown and cannel coals, lignite, peat, or turf, etc.), are the results of a slow decay, under certain conditions, of vegetable matter. The proper conditions for its formation seem to be, enough water to cover the fallen timber or vegetation, and then a covering of clay or mud, to exclude the air and apply pressure to the mass. The liquid and volatile portions are gradually lost, leaving the carbon behind. The final product of this change, assisted by subterranean heat, is **anthracite** coal, which often contains 96 to 98 per cent. of carbon, and almost no volatile products. Petroleum oil is believed to be the expelled liquid portions of the wood, separated from the anthracite coal by heat and pressure. **Bituminous** coal is a softer, less compact, less decayed variety. It often contains the structure of the wood. It burns with a smoky flame, while the anthracite furnishes very little or no flame. **Brown coal, lignite,** and **wood coal,** are names of less perfect varieties than those before mentioned. **Peat,** or **turf,** is a mixture of mud with partially decayed plants and roots, obtained from certain marshy districts. It is used as fuel. **Cannel coal** is a compact, even-textured coal, without lustre, and takes fire readily, burning with a clear, yellow flame. It has been used as candles—hence the name.

274. **Properties.**—Carbon is insoluble in all ordinary menstrua, but soluble to a slight extent in molten cast-iron, forming a carbide. It is fused and volatilized only in the voltaic arc. At ordinary temperatures it is permanent in air and all chemical reagents. At high temperatures it has a strong affinity for oxygen, and on this account it is used as a reducing agent in smelting the ores, from which it removes the oxygen. Indi-

rectly, it enters into combination with a great many of the elements. In combination with hydrogen, nitrogen and oxygen, it forms a large share of organic bodies. Most combustible bodies contain carbon.

275. **Coal Gas.**—Illuminating gas, as it is often called, is made on a large scale by the dry or destructive distillation of bituminous coal. In principle, the manufacture is simple; but in practice, it requires considerable skill to prepare a good illuminating gas. During the distillation, which is conducted in horizontal, semi-cylindrical, fire-clay or cast-iron retorts set in brickwork, a variety of products are produced besides the gas, such as tar, heavy oils, lighter oils, steam, ammonia from the nitrogen of the coal, etc. There is left in the retort a porous, friable mass, called **coke**. After the retorts have been used for some weeks, there is to be found lining their inner surfaces a very compact layer of carbon, usually known as **gas retort carbon**, and used in making the negative plates in some forms of electric batteries, and the poles of electric lights. In many of its properties it resembles graphite.

The coal is distilled at a bright red heat, and the volatilized products are conducted into a large, horizontal, iron pipe, half filled with water, into which the pipes from the retorts dip. This is called the **hydraulic main**. A large part of the **coal tar** and heavy oils condense in this main. The volatile products are then made to traverse a series of upright tubes, in the form of an inverted U, called condensers. The lower end of one limb of each condenser dips under water, so as to cool the gas, and to dissolve the ammonia in it. After traversing a number of these condensers, where the remainder of the tar, steam and ammonia liquor are condensed, the gas passes to the purifiers, which are composed of a series of large boxes, in which are several perforated shelves, or trays, holding fresh slaked lime, or a mixture of sawdust and iron oxide. The gas is passed slowly through these purifiers, so as to expose it to the action of the lime or iron oxide, to remove H_2S, CO_2 and other volatile acids, if present. In some plants there is an additional process of washing the gas by passing it through weak sulphuric acid, to remove the small quantity of ammonia still remaining. The tar and ammonia liquor are sold as by-products, and the latter furnishes a very considerable portion of the ammonium compounds of the market. After passing through the purifiers, the gas is conducted in underground pipes to the gasometer, to be stored until needed.

The **gasometer** is a very large, tub-shaped vessel, made of boiler-iron floated bottom upward upon water, and balanced by weights attached to chains passing over pulleys at the top of iron pillars, which are erected around the gasometer for that purpose. As the gas is forced into the gasometer, the latter rises out of the water, and sinks again as the gas is used.

Various other processes for the manufacture of gas have been used with varying success; such as, the distillation of the heavy petroleum oils, either alone, with coal, or with admixture of the vapors with air or steam.

Gasoline, or air gas, is air saturated with the vapors of the very volatile oils from petroleum.

Water Gas is very largely used in all large cities, owing to cheapness of manufacture. It is made by the action of steam upon coal heated to a red or white heat. The gas thus produced, is, in some works, mixed with vapor from naphtha, and then again strongly heated in retorts, and finally purified, as in ordinary coal gas.

Composition.—The composition of coal gas varies somewhat with the composition of the coal used, and the temperature of the retorts during the distillation.

The following figures represent the composition of coal gas and water gas supplied to Brooklyn in 1883:—

	Coal Gas.	Water Gas.
Carbon dioxide	0.0	0.3
Carbon monoxide	7.9	28.25
Hydrogen	50.2	30.3
Illuminants (C_2H_4, C_2H_2, etc.)	4.3	12.85
Marsh gas (CH_4)	29.8	21.45
Nitrogen	7.8	6.85
	100.00	100.00

(*Remsen*).

276. **Carbon and Oxygen.**—There are four oxides of carbon known, having the formulæ C_8O_3, C_4O_3, CO and CO_2. The first of these compounds is a light brown powder, formed by heating the suboxide (C_4O_3).

Carbon Suboxide is an amorphous extractive matter formed by the action of an electric current upon carbon oxide.

Carbon Monoxide (CO–) is a colorless, tasteless, almost odorless, combustible, poisonous gas.

It is always produced when carbon or bodies containing it are burned with an insufficient supply of oxygen or air, or by conducting carbon dioxide over or through red hot coals, $CO_2 + C = 2CO$. It may be prepared by warming oxalic acid with sulphuric acid.

$$H_2C_2O_4 + H_2SO_4 = H_4SO_5 \text{ (or } H_2O, H_2SO_4) + CO_2 + CO.$$

The mixed CO_2 and CO are passed through a solution of sodium hydrate to absorb the CO_2, and the CO remains. One part of potassium ferrocyanide warmed with nine parts of H_2SO_4 may be used to give the gas in nearly a pure state. Its density is 14. It is almost insoluble in water, but soluble in a solution of cuprous chloride in ammonium hydrate. It burns in the air with a bluish-lavender flame, producing the higher oxide —CO_2. Owing to this property, it plays an important part in the reduction of ores in the blast furnace. It does not support combustion or respiration. It diffuses readily through red hot cast-iron, and frequently escapes from stoves and hot air furnaces.

277. **Physiological Effects.**—It has the power of combining with the hæmoglobin of the blood and expelling the oxygen. It thus acts as a narcotic poison, causing dizziness, headache, nausea, incoördination of movements, convulsions, and death. If the carbon monoxide be in sufficient quantity to saturate all the hæmoglobin, recovery is seldom, if ever, realized. The blood has a light red color, and does not coagulate or decompose as readily as normal blood. If the hæmoglobin is only partially saturated, recovery may take place, but very slowly—debility, anorexia, etc., remaining for days. Air containing 0.5 per cent. kills birds in three minutes; 2 per cent. rendered a guinea pig insensible in two minutes. Artificial respiration is of little use. Transfusion of blood is the most promising treatment. The sources of danger are open fires, defective draught in chimneys, escape of coal gas, and especially "water gas," from defective fittings or from leaks under the ground. When the ground is frozen and the gas escapes into the soil, near a cellar, the gas diffuses through the ground into the cellar, and, as it is thus deprived of its odor, persons may be poisoned and not know where it comes from. Coal gas poisoning is essentially a poisoning by the carbon monoxide which it contains. Suffocation by coal gas is not very different from suffocation by other gases.

278. **Carbon Dioxide—Carbonic Anhydride** (CO_2), sometimes called carbonic acid gas, is found free in the air in the proportion of about 4 parts per 10,000; and in ordinary well ventilated rooms from 5 to 6 parts per 10,000. It is found in volcanic gases, and in solution in many mineral springs. It sometimes accumulates in mines, wells and cellars, in dangerous quantities, and is then known as "choke damp." It may be detected in such places by lowering a candle. (See below.) It is produced when carbon or its compounds are burned with a free supply of air, by alcoholic and other fermentations, by the respi-

ration of animals, and by slow oxidation of organic matter in the natural process of decay. In the laboratory it is obtained by the action of an acid upon a carbonate.

$$CaCO_3 + 2HCl = CO_2 + CaCl_2 + H_2O.$$

Carbon dioxide at ordinary temperatures and pressures is a colorless, transparent, odorless, tasteless, (by some sweetish), gas. Sp. gr. = 1.524. Density = 22. Under a pressure of 36 atmospheres at 0°C.(32°F.), it is condensed to a colorless, mobile liquid, of sp. gr. 0.94. Above 32.5°C. it cannot be liquefied at any pressure. This is known as the **critical point in temperature.** When the liquid is exposed to the air it rapidly evaporates, producing a temperature so low as to freeze a portion of it to a snow-like solid; the temperature being sometimes as low as —130°C. (—202°F.). The gas extinguishes the combustion of burning bodies, and animals die very quickly in it. Death has resulted from persons entering mines, wells and fermenting vats, where the gas had accumulated. It is unsafe for a man to venture into a well or other place where a candle will not burn. CO_2 is soluble in its own volume of water, at the ordinary temperature and pressure, forming a solution of carbonic acid (H_2CO_3). Common soda water is a solution of the gas in water under pressure; it contains no sodium salt, as its name would imply.

279. **Physiological Effects.**—These vary with the degree of concentration of the gas and its dilution with other gases. If the gas be pure, it causes death instantly, by apnœa from spasm of the glottis. When somewhat diluted, there is, at first, great loss of muscular power; the patient becomes livid, sinks down and dies without a struggle. When still more dilute, there is, at first, irritation of the throat, then giddiness, ringing in the ears, loss of muscular power, with rapid pulse and respiration, and occasionally vomiting and convulsions, finally ending in coma and death.

The amount of the gas that can be tolerated in the air depends not only upon the quantity of it actually present, but also upon the source of it. Thus, when the source of the gas is animal respiration or combustion, the oxygen is withdrawn from the air at the same time, and a much smaller quantity will prove fatal than where the gas is simply added to the normal atmosphere.

If the CO_2 is simply added to the air, 10 per cent. may be regarded as poisonous, and even 8 per cent. will prove injurious. If, on the other hand, the oxygen be increased, an air containing even 20 per cent. may be breathed. A taper will burn in an air containing 8 per cent. of CO_2, provided the oxygen be present in normal quantity, and will burn feebly in such an air containing 10 per cent. Where the CO_2 is produced by combustion or respiration, the injurious effects are soon perceived, and are due to several causes; viz., the deficiency of oxygen, the presence of too great a quantity of CO_2 and moisture, the rise in temperature, and the action of the organic matter exhaled from the lungs and skin.

The exposed air contains from 4 to 5 per cent., or about .78 cubic feet of CO_2 per hour, and there is absorbed .94 cubic feet of oxygen. A stearin candle gives off .5 cu. ft. of CO_2, and uses up 1 cu. ft. of oxygen. A gas light burning 5 feet of gas per hour (12 candle power), gives off very nearly 6 cu. ft. of CO_2, or 3.7 times as much as one man; as much heat as 2 men; removes more oxygen than 5 men; and gives off nearly as much watery vapor as 5 men. More than 6 parts of CO_2 per 10,000 of air renders it oppressive, and should not be allowed. Assuming the amount of CO_2 given off in an hour by an adult to be .7 cu. ft., and normal air to contain 4 parts, it would require about 3500 cu. ft. of air per hour, in order that it should not receive more than two parts per 10,000 of CO_2, or 2 cubic feet. Dr. Parkes fixes the amount necessary at 2000 cu. ft. per hour. It is impossible to change the air of a room oftener than 3 or 4 times per hour without causing uncomfortable draughts; and it would, therefore, require 700 to 1000 cu. ft. of room space, in order to keep the air of the room in a proper condition. If lights are used, which change the air, a corresponding calculation must be made for them. A common oil lamp, or two sperm candles (not an argand lamp), will contaminate the air about the same as an adult man. The English Poor-Law Board's requirements for dormitories, to prevent over-crowding, are—

Cu. ft.	
1200	Lying-in cases and offensive sick.
850	Sick.
700	Infirm. Same room night and day.
500	Infirm. Separate room in day.
300	Healthy.

These figures are, of course, too small for general use. Parkes quotes Morin as giving the following amount of fresh air necessary to be furnished to each adult per hour in the following circumstances:—

	Day.	Night.
In Barracks,	1059 cu. ft. per hour.	2118 cu. ft. per hour.
" Workshops,	2118 " " " "	
" Schools,	1059 " " " "	
" Hospitals,	2825 " " " "	

In sleeping apartments, the amount of space allowed to each individual should be not less than 1000 feet, *i. e.*, a room 10 × 10 × 10 ft.; but, as the proportion of carbonic dioxide would accumulate slowly, and as a much larger amount of the gas may be borne without serious discomfort, even one-half this capacity may be tolerated with little inconvenience beyond a feeling of fatigue or sleepiness in the morning. Dr. Tidy regards 400 cubic feet as the very smallest amount of space that should be allowed to each person in a sleeping room, to avoid serious overcrowding. The foulest air in an occupied room is at the ceiling. The heat of the body or a lamp causes an expansion of the air about them and an upward current of air, the heated CO_2, water vapor, etc. These gases reach the ceiling before they cool sufficiently to stop this upward current, and before there is time for perfect diffusion. The upper galleries in theatres are supplied with impure air from the main floor and lower galleries, and from the gas lights. Fresh air should always be admitted to a room near the floor, and the outlet for impure air should be at or near the ceiling. It must be remembered that the law of diffusion of gases does not allow the CO_2 to accumulate in one part of the room and remain there for any considerable time, but mixes them evenly.

Nor is this diffusion confined within a room. It takes place through porous walls between the indoor and outdoor air, especially in winter, when there is much difference in the temperature of the two. Indeed, a very fair amount of ventilation may be effected in this way, where the walls are of brick or stone, and not painted or papered inside.

Carbonic dioxide exists in the blood in solution, and does not combine with the corpuscles, the blood after death remaining dark-colored and liquid. Putrefaction after CO_2 poisoning is slow, while animal heat and rigidity are very persistent.

280. **Tests for Carbon Dioxide.**—If the quantity exceeds 12 per cent., a taper is extinguished. Moistened litmus paper is first reddened, then bleached. Lime water or baryta water is rendered cloudy by it, from the precipitation of the carbonates of calcium or barium.

281. **Carbonic Acid and Carbonates.**—CO_2 is soluble in water, with which it combines to form carbonic acid (H_2CO_3).

$$H_2O + CO_2 = H_2CO_3.$$

Carbonic acid is a feeble, dibasic acid, forming a double series of salts, the **carbonates** and **acid** or **bi-carbonates.** As mentioned above, "soda water" is a solution of carbonic acid in water, kept under pressure. When the pressure is removed, a large portion of the acid undergoes decomposition into the anhydride (CO_2) and water. The same decomposition, with effervescence, takes place when we prepare the acid by treating a carbonate with a stronger acid. The carbonates of gold, arsenic, antimony and aluminum are unknown. The carbonates of the alkaline metals are soluble in water, and are not decomposed by heat; while the carbonates of all the other metals are insoluble and decomposed by heat into the oxides of the metals. The bicarbonates are formed by passing carbonic dioxide through the solution of the carbonates. They are converted into the carbonates by heat. The carbonates of ammonium are volatilized by heat. Water charged with carbonic acid dissolves the carbonates of some of the metals, as calcium, magnesium, iron, copper, lead, etc., which gives rise, in the case of the first two, to hard water. This hardness is deposited again on boiling the solution, or even on free exposure to the air, thus forming an important element in certain geological formations.

282. **Carbon and Sulphur—Carbon Disulphide** (CS_2), is formed, like the dioxide, by the direct union of the elements. When the vapor of sulphur is passed over charcoal heated to redness, the two elements combine, producing vapors of carbon disulphide, which condense into a very volatile, colorless, mobile liquid, possessing a peculiar, disagreeable odor; it refracts light strongly, and for this reason is used to fill hollow glass prisms for the spectroscope. It is combustible, burning with a blue flame, and has been suggested as a means of furnishing sulphurous oxide for fumigation, by burning a mixture of CS_2 and alcohol in a lamp.

The vapor mixed with air, forms an explosive mixture, and mixed with nitrous oxide, it burns with a very brilliant flame. It is insoluble in water, but is miscible with alcohol and ether. It is a ready solvent for sulphur, phosphorus, caoutchouc (India rubber), fatty oils and iodine, with the last of which it forms a violet-red solution. CS_2 dissolves in a solution of the alkaline sulphides forming **sulpho-carbonates,** $CS_2 + K_2S = K_2CS_3$. CS_2 may be regarded as the anhydride of sulpho-carbonic acid, H_2CS_3, obtained by adding hydrochloric acid to a sulpho-carbonate. **Carbon Monosulphide** (CS)—a brown-red powder, and a sulphide having the formula C_5S_2, and an oxy-sulphide, COS, are also known.

290. **Tests.**—1st. Silver nitrate precipitates the acid as silver cyanide—a white, curdy precipitate, insoluble in nitric acid. A rod moistened with $AgNO_3$ and held in the vapor, is rendered milky. 2d. Add a solution of **liquor potassæ**, or potassium hydrate, then a solution of ferrous sulphate mixed with ferric sulphate, then a small quantity of sulphuric acid, when a blue deposit of Prussian blue will appear. 3d. To a small portion of the suspected liquid, in a wide test tube or crucible, add dilute sulphuric or hydrochloric acid. Invert over this a watch glass, convex side down, with one or two drops of yellow ammonium sulphide upon the under side of it. Warm the crucible gently, and after a few minutes remove the watch glass, warm, and evaporate the ammonium sulphide by blowing upon it. Now touch the stain with a drop of ferric chloride, when a blood-red stain will make its appearance, due to the formation of sulphocyanate of ammonium on the glass, which, with the iron, strikes a red color. This is easy to perform, is applicable to organic mixtures, and is quite delicate.

291. **Cyanic Acid** (H–O–Cy).—Metallic cyanides readily take up oxygen, when fused with oxidizing agents, such as potassium nitrate, or even the oxides of some of the metals, and form cyanates. The acid, HOCy, may be prepared by decomposing its salts with dilute acids, but it is unstable, and breaks up into carbon dioxide and ammonia.

$$HOCN + H_2O = NH_3 + CO_2.$$

Of the cyanates, the ammonium cyanate is the most interesting, as its solution in water, on being heated, forms urea, a body isomeric with it, and a well known excretory substance found in urine.

$$NH_4CNO = CO(NH_2)_2.$$

This is interesting as being the first animal substance prepared by synthesis.

292. **Sulphocyanates.**—The potassium salt is prepared by fusing potassium ferrocyanide and sulphur together, and exhausting the fused mass with alcohol. On evaporating the alcohol, a white, crystalline salt is obtained, soluble in water, and having the formula KSCy.

Ammonium Sulphocyanate (NH_4SCy) is prepared by heating hydrocyanic acid with yellow ammonium sulphide.

The principal interest attached to these salts is as tests for ferric iron salts, with which they give a blood-red solution.

Mercuric sulphocyanate, formed by precipitating mercuric

potassium sulphocyanate, is decomposed by heat, the
up and leaving a voluminous residue. It is used
toy, **Pharaoh's Serpents.**

pound Cyanides.—Cyanogen shows a remarkable
form complex compounds. Among these more
mpounds are two series of bodies in which cyanogen
form the radical. Of these, the **ferrocyanides** and
anides of potassium and iron will be mentioned here.

assium Ferrocyanide—$K_4(FeCy_6)$—**Yellow Prus-**
of **Potash,** is an important commercial product, manu-
ured on a large scale by fusing in a closed crucible
rogenous animal matter (horns, hoofs, leather scraps, etc.),
ith potassium carbonate and iron filings, treating the fused mass
with water, and crystallizing. The salt is thus obtained in large, yellow, tabular crystals. It is used in dyeing, in preparing certain pigments, and as a source of all the other cyanogen compounds. By simple fusion, potassium cyanide (KCy) is prepared.

$$K_4FeCy_6 = 4KCy + FeC_2 + N_2.$$

From a solution of the salt, various other ferrocyanides are prepared by precipitation. Of these, **Prussian Blue** (ferric ferrocyanide), prepared by adding to a solution of potassium ferrocyanide a solution of some ferric salt, is used as a pigment and as a medicine. With a ferrous salt the precipitate is white, but quickly becomes blue by oxidation. This test serves to distinguish ferrous from ferric salts. With cupric salts we obtain a reddish-brown precipitate, Cu_2FeCy_6.

Ferricyanides.—By passing chlorine through a solution of K_4FeCy_6, a compound is formed in which the iron of the radical is tetrad, and the radical itself becomes hexad. $K_6\,(Fe_2Cy_{12})^{VI}$. On evaporating, we obtain dark red crystals of **potassium ferricyanide,** or **red prussiate of potash.** With ferrous salts a solution of this salt gives a deep blue precipitate of ferrous ferricyanide—$Fe_3Fe_2Cy_{12}$—Turnbull's Blue. Ferric salts give no precipitate or coloration, and thus we distinguish ferric from ferrous salts.

Ferric Ferrous cyanide = Prussian Blue.
Ferrous Ferric cyanide = Turnbull's Blue.

The further consideration of carbon compounds will be considered in part third.

THE ALKALINE METALS.

1. Lithium	7	4. Rubidium	85.5
2. Sodium	23	5. Cæsium	132.6
3. Potassium	39	6. (Ammonium)	

294. The metals of this group present great similarity in their chemical and physical properties. Exposed to the air, they all oxidize readily. They decompose water violently, with the formation of strong basic hydrates which dissolve in the excess of water. The hydrates thus formed are called **caustic alkalies** (caustic potash, caustic soda); hence the name, alkaline metal. Nearly all of the salts of these metals are soluble, and most of them, when in solution, turn blue litmus red. They form but one chloride, one iodide and one bromide.

LITHIUM.

Li = 7.

295. Lithium occurs widely distributed in nature, but in small quantities. It is found in some mineral springs and in the ashes of many plants, chiefly that of tobacco and beet. It is usually obtained by separating it from its chloride by electrolysis. It is silver-white in color, and decomposes water at ordinary temperatures. It is the lightest of the solid elements,—sp. gr. 0.589,—and floats upon naphtha. It fuses at 180°C. (356°F.), and burns in air with an intense red light. Its salts closely resemble those of sodium.

Lithium Oxide (Li_2O) is a white solid, formed by burning lithium in dry oxygen. It slowly dissolves in water, forming the hydrate, LiOH.

Lithium Chloride (LiCl) crystallizes at ordinary temperatures, in regular, anhydrous octahedra; below 10°C. (50°F.), however, with two molecules of water. It is very deliquescent.

296. **Lithium Bromide** (LiBr) is obtained by decomposing lithium sulphate with potassium bromide; or, by neutralizing a solution of hydrobromic acid with lithium carbonate. It crystallizes in deliquescent needles.

297. **Lithium Carbonate—Lithium Carbonas** (U. S.), (Li_2CO_3), is obtained by fusing a native silicate called **lepidolite**, with barium sulphate and carbonate, and potassium sulphate. It is then extracted with water, and precipitated with sodium carbonate. It is a white, odorless powder, of a strongly alkaline taste, soluble in water to the extent of 12 parts per 100. It unites readily with uric acid, forming a soluble lithium urate. It

is said that 250 parts of lithium carbonate at a temperature of 38°C. (100.4°F.), will dissolve almost 1000 parts of uric acid. This property renders it of value in diminishing the deposit of uric acid formed in gout, and in dissolving uric acid calculi.

SODIUM—Natrium.

Na = 23.

298. **Occurrence.**—This metal occurs widely distributed, being found in sea-water and in rock salt as chloride, and in many native silicates.

299. **Preparation.**—Like potassium, it is obtained by a process which depends upon the reduction of the carbonate by coal. The two substances are mixed in an iron retort, and heated to redness.

$$Na_2CO_3 + 2C = 2Na + 3CO.$$

300. **Properties.**—Sodium is a soft, silver-white metal, resembling potassium, but less easily oxidized. It becomes slowly coated with a brownish-yellow layer on exposure to the air, and should be kept under naphtha. It fuses at 95.6°C. (203.8°F.), and volatilizes at a white heat, the vapor burning with a bright yellow flame. Sp. gr. 2.078. It is characterized by its affinity for oxygen, decomposing water at ordinary temperatures, liberating hydrogen, and forming sodium hydrate.

301. **Oxides.**—Three are described—Na_2O, Na_4O, Na_2O_2. The first of these is a white powder, formed by the oxidation of the metal in dry air. It is very deliquescent, soon liquefying in air. The peroxide is a grayish-white mass, obtained by burning sodium in a current of oxygen. They all unite with water with great energy.

302. **Sodium Hydrate—Caustic Soda—Soda** (U. S., Br.), ($NaOH$), is usually obtained by boiling a solution of sodium carbonate with calcium hydrate.

$$Na_2CO_3 + Ca\,(OH)_2 = CaCO_3 + 2NaOH.$$

The resulting solution, after filtering, is evaporated to dryness, dissolved in alcohol, again evaporated, fused in a silver vessel, and cast into sticks. This product is usually labeled **caustic soda by alcohol.** It is a white, opaque, brittle, crystalline mass, fusing below redness; sp. gr. 2.00. It dissolves readily in water, the solution being known as **soda lye** in the arts, and in pharmacy, as **liquor sodæ** (sp. gr. of latter 1.071). This solution attacks glass; hence, the necks and stoppers of bottles containing it should be coated with paraffine. When exposed to the air, sodium hydrate attracts water and carbonic dioxide, and is converted into the carbonate.

303. **Sodium Chloride—Common Salt—Sodii Chloridum** (U. S., Br.), ($NaCl$), is found very abundant in nature. It is deposited in almost all parts of the globe in the solid form as rock salt; in solution, it is found in all natural waters, and to the extent of 2.7 to 3.2 per cent. in sea water; it also exists in most animal and vegetable tissues. It is formed in a great number of chemical reactions. Its most important source is the deposits of rock salt, from which it is mined; it is also obtained by the evaporation of sea or saline spring waters. It crystallizes from water in translucent cubes. Sp. gr. 2.078. It fuses at a red heat, and volatilizes at a white heat. Hot water dissolves but little more than cold; 100 parts of water at 0°C. (32°F.) dissolve 36 parts of the salt; and at 100°C. (212°F.), 39 parts. A saturated solution, therefore, contains 26 per cent. sodium chloride. From dilute solutions nearly pure ice is obtained by freezing. On account of a slight admixture of magnesium salts, most specimens of common salt will deliquesce; the perfectly pure salt, however, is not hygroscopic.

304. **Sodium Bromide** ($Na\,Br$) is formed by the action of bromine upon a solution of sodium hydrate. It crystallizes in anhydrous cubes, and is soluble in 1.13 parts of water at 20°C. (68°F.), and in 0.87 at 100°C. (212°F.) It contains 77.67 per cent. of bromine.

305. **Sodium Iodide** (NaI) also crystallizes in cubes without water; and is soluble in 0.56 parts of water at 20°C. (68°F.), and in 0.32 at 100°C. (212°F.) It contains 84.66 per cent. iodine.

306. **Sodium Sulphate—Neutral Sodium Sulphate—Glauber's Salt Sodii Sulphas** (U. S.)—**Sodæ Sulphas** (Br.)—(Na_2SO_4), occurs native in deposits; also in solution in mineral waters. It is a by-product in the manufacture of sodium chloride from sea water and brine, and in several manufacturing industries. It is prepared by the decomposition of common salt with sulphuric acid.

$$2NaCl + H_2SO_4 = Na_2SO_4 + 2HCl.$$

Sodium sulphate crystallizes, at ordinary temperatures, with two proportions of water (7 aq. and 10 aq.) in large, colorless, monoclinic prisms, which effloresce in the air, losing all of their water. If heated to 33°C. (91.4°F.) it fuses in its own water of crystallization, and at higher temperatures becomes anhydrous.

The following curious action of the solution of Glauber's salt may also be noticed: If the solution, saturated at 33°C. (91.4°F.), be cooled down to the ordinary temperature, and even far below, no separation of crystals occurs, although the salt is very much more insoluble at lower temperatures than at 33°C. (91.4°F.). This formation of a **supersaturated solution** is common to many salts, though not to so marked a degree as in the case of Glauber's salt. This supersaturated solution may be agitated, and still no crystals form. But, if it be gently touched with a glass rod or some other solid body, the entire mass will at once become crystallized.

307. **Hydro-Sodium Sulphate—Acid Sodium Sulphate—Sodium Bisulphate** ($NaHSO_4$), is obtained by the action of sulphuric acid upon sodium sulphate, or sodium chloride.

$$NaCl + H_2SO_4 = NaHSO_4 + HCl.$$

It crystallizes in long, four-sided prisms. It fuses readily, and at higher temperatures loses water and is converted into the pyrosulphate $Na_2S_2O_7$. Very soluble in water, giving an acid solution.

308. **Sodium Hyposulphite—Sodium Thiosulphate** ($Na_2S_2O_3$), is prepared by boiling the sulphite with sulphur.

$$Na_2SO_3 + S = Na_2S_2O_3.$$

It forms large monoclinic prisms, which contain 5 molecules of water, and is slightly deliquescent in the air. It is used as a reducing agent, decoloring an iodine solution with the formation of sulphuric acid and sodium iodide.

309. **Sodium Carbonate—Soda—Neutral Carbonate of Soda—Sal Soda—Washing Soda—Sodii Carbonas** (U.S.) **—Sodæ Carbonas** (Br.), (Na_2CO_3), is the most important of the sodium compounds, for industrial purposes. It occurs abundantly in nature, in the so-called sodium seas (in Egypt, and the Caspian Sea), and is contained in the ashes of many sea plants, chiefly the algæ. The principal supply is sodium chloride, from which it is manufactured according to a method devised by Leblanc, in 1808. By this method the sodium chloride is first converted into sulphate by warming with sulphuric acid. The sulphate, when dried, is mixed with charcoal and calcium carbonate, and strongly heated. Two reactions take place during this process: **First,** the carbon reduces the sodium sulphate to sulphide.

$$Na_2SO_4 + 2C = Na_2S + 2CO_2.$$

Second, the sodium sulphide and calcium carbonate react to form calcium sulphide and sodium carbonate.

$$Na_2S + CaCO_3 = CaS + Na_2CO_3.$$

The high temperature also converts a portion of the calcium carbonate into calcium oxide and carbon dioxide. The products of this fusion, known as **black ball soda**, are, therefore, sodium carbonate, calcium sulphide and calcium oxide. The black ball is broken up and lixiviated with hot water, which dissolves out the sodium carbonate; this solution is evaporated to dryness, and crude **soda ash**, or **soda**, of commerce results.

Of late years another process, known as Solvay's, or the ammonia method, has largely replaced that of Leblanc. In this process, ammonium bicarbonate reacts upon sodium chloride, forming the sparingly soluble sodium bicarbonate, and the freely soluble ammonium chloride. The sodium bicarbonate is then converted into the carbonate by heat. At ordinary temperatures, sodium carbonate crystallizes in large, rhombic crystals, containing 10 molecules of water, which effloresce in dry air. It is soluble in water, most freely at 38°C. (100.4°F.) Its solutions have an alkaline reaction. When the crystals are calcined at a dull red heat, they disintegrate, give off their water of crystallization, and form a white powder, the **sodii carbonas exsiccata**, of the U. S. P. ·100 parts of H_2O dissolve 10 parts of this anhydrous carbonate at 0°C. (32°F.), and 138 parts at 38°C. (100.4°F.).

310. **Hydrogen Sodium Carbonate—Acid Sodium Carbonate—Sodium Bicarbonate—Sodii Bicarbonas** (U. S.)—**Sodæ Bicarbonas** (Br.), ($NaHCO_3$), is found in many mineral waters. It is produced by the ammonia process as above, and by the action of carbon dioxide upon sodium carbonate.

$$Na_2CO_3 + CO_2 + H_2O = 2NaHCO_3.$$

It forms small, rectangular prisms, which are anhydrous, but dissolve in ten or eleven parts of water. Its solutions are nearly neutral to test paper. By heating the solid or boiling its solutions, it gives off carbon dioxide and is converted into the carbonate.

311. **Sodium Phosphates.**—These are three in number. They are less soluble and crystallize better than the potassium salts of phosphoric acid.

The **tri**-sodium phosphate, or **basic** phosphate (Na_3PO_4), is made by saturating 1 part of phosphoric acid with 3 parts of sodium hydrate. It crystallizes in six-sided prisms, and is soluble in 5.1 parts of water at 15.5°C. (59.9°F.) Its solution is alkaline to test paper.

Hydrogen Disodium Phosphate—Disodium Phosphate—Neutral Sodium Phosphate—Sodii Phosphas (U. S.)—**Sodæ Phosphas** (Br.), (Na_2HPO_4), is more stable than the other phosphates, and is the one generally employed in medicine and in laboratories. It may be prepared by treating phosphoric acid with sodium hydrate to feeble alkaline reaction. Below 30°C. (86°F.) it crystallizes in large, rhombic prisms, with 12 Aq.; at 33°C. (91.4°F.) it crystallizes with 7 Aq. The salt with 12 Aq. effloresces in air, losing 5 Aq.; that with 7 Aq. does not. Both are freely soluble in water, and show a faintly alkaline reaction.

Monosodium Phosphate—Acid Sodium Phosphate (NaH_2PO_4), crystallizes in rhombic prisms with 1 molecule of water, and is acid in reaction. At 100°C. (212°F.) it loses its water of crystallization, and at about 200°C. (392°F.) forms **sodium pyrophosphate** ($Na_2H_2P_2O_7$), and at 240°C. (464°F.) the **metaphosphate** ($NaPO_3$).

$$Na_2H_2P_2O_7 = 2NaPO_3 + H_2O.$$

312. **Sodium Nitrate—Chili Saltpetre—Sodii Nitras** (U. S.), ($NaNO_3$), is found native in extensive deposits in Peru. It crystallizes in rhombohedra, which closely resemble cubes; hence, it is called cubic saltpetre. It is deliquescent, and is, therefore, not adapted for the manufacture of gunpowder; it has a cooling, saline, somewhat bitter taste. It is more readily soluble in water than potassium nitrate, which, in other respects, it quite closely resembles (which see). It is used in the manufacture of nitric acid, and as a fertilizer.

313. **Sodium Borates.**—Six are known. The only one of importance is the **Disodium Tetraborate—Sodium Pyroborate—Borax—Tincal—Sodii Boras** (U. S.), ($Na_2B_4O_7$), which is found native in some of the lakes of Thibet, from which country it was formerly imported.

At present it is prepared artificially by boiling boric acid with sodium carbonate. Boric acid is found in lagoons of Tuscany, and this is the present source. It crystallizes in large, hexagonal prisms, with 10 molecules of H_2O, or in regular octahedra with $5H_2O$. The former variety effloresces in dry air; the latter is permanent. Both dissolve in 12 parts of cold and 2 parts of boiling water, forming a solution that has a feebly alkaline reaction. Upon heating, both salts puff up considerably, lose their water, and form a white, porous mass (**burned borax**), which finally fuses to a transparent glass. In the fused state, it will dissolve many metallic oxides, forming clear glasses (borax beads), which often show a characteristic color on cooling; thus, copper salts give a blue, and chromic oxide an emerald green glass. Borax is used in this way as a blowpipe test for certain metals.

It is this property of dissolving oxides of the metals that renders borax useful in welding and soldering metals. In these operations it is used to remove the oxide, or rust, from the surfaces of the metals.

314. **Sodium Hypochlorite** ($NaClO$).—When in solution **Liq. Sodæ Chloratæ** (U. S.), (Br.)—Labarraque's solution. It may be prepared by decomposing a solution of chloride of lime with sodium carbonate. It yields up its chlorine readily, thus acting as an efficient disinfecting and decolorizing agent.

315. **Physiological Effects of the Sodium and Potassium Compounds.**—The action of the halogen salts of these metals is generally that of the combined chlorine, bromine, or iodine. The hydrates of both metals, and, to a lesser degree, the carbonates, tend to disintegrate the tissues with which they come in contact; hence, they possess powerful caustic properties. If taken internally, the hydrates are highly poisonous, causing death, like the mineral acids, either immediately, by their corrosive properties, or secondarily, by exciting inflammation of the gastro-intestinal mucous membrane, with consequent thickening and constriction.

In cases of poisoning by the caustic alkalies, the stomach should be evacuated, and a weak acid, such as dilute vinegar or lemon juice, given, to neutralize the alkali; or, it should be saponified by the administration of some oil or fat. The nitrate of these metals is also toxic in its influence, and for it there is no direct antidote. The alkaline carbonates are, undoubtedly, of considerable importance to the carrying on of the normal functions of the animal body. In the first place, it is exceedingly probable that some, at least, of the albuminoid matters of the blood are held in solution by reason of its alkaline reaction, which is largely given it by these carbonates.

Secondly, it has been shown very clearly that the alkaline reaction of the blood is of first importance to the oxidation processes, which are intimately connected with the production of animal heat and retrograde metamorphosis. It is only in the presence of a free alkali that many organic substances will unite with oxygen, and thus, their decomposition without an alkali would be impossible at the temperature of the body. In proof of this, it is known that, if the vegetable acids are given free, they will reappear, for the most part, unchanged, in the urine; but if they are in combination with the alkalies when given, they are thoroughly burned up in the blood, and reappear as carbonates. In fact, so important are these alkaline salts—carbonates and phosphates—that without them, albuminoid bodies will not support life.

The alkaline carbonates in sufficient quantity render the urine alkaline in reaction and increase the quantity. The tartrates, citrates and acetates of sodium and potassium have a very similar action upon the economy to that of the carbonates, into which they are converted either in the intestines or blood. A slightly more cathartic action is attributed to the tartrates than is possessed by the carbonates. This action is also more or less shown by the sulphates and phosphates.

POTASSIUM (Kalium).

$K = 39.$

316. **Occurrence.**—This metal is found widely distributed in rocks and minerals, principally as silicates. By the action of the atmosphere and other influences, these silicates gradually decompose, the potassium passes into the soil and is absorbed by the plants, from the ashes of which it may be obtained. The chloride and sulphate are also found in sea-water, and in large deposits, mixed with other chlorides. A chloride of potassium is mined in Stassfurth, Germany, as a source of potassium salts.

317. **Preparation and Properties.**—It is prepared by calcining an intimate mixture of the carbonate with carbon.

$$K_2CO_3 + 2C = 2K + 3CO.$$

Such a mixture may be made by heating organic potassium salts, as crude tartar, to redness. In this way a black mass is formed, consisting of potassium carbonate and free carbon. By heating this black mass to a white heat in an iron retort, the potassium distills off, and is condensed under mineral naphtha. Potassium is a silver-white, lustrous metal, brittle at 0°C. (32°F.), waxy at 15°C. (59°F.), fuses at 62°C. (143.6°F.), and distills at a red heat. Sp. gr. at 15°C. (59°F.) = 0.865. Its affinity for oxygen is such that, if it be exposed to the air, it tarnishes at once. It decomposes water or ice with great energy, with the formation of potassium hydrate and the liberation of hydrogen, which is ignited by the high temperature caused by the reaction. It combines directly and energetically with the halogens, sulphur, phosphorus, arsenic, antimony and tin.

318. **Potassium with Oxygen—Potassium Oxide** (K_2O), results from the direct oxidation of potassium, by simply exposing thin strips of the metal to dry air, or by the action of potassium upon the hydrates.

$$2KOH + K_2 = 2K_2O + H_2.$$

It is a white, deliquescent, caustic powder, uniting readily with water to form the hydrate, with the evolution of much heat.

319. **Potassium Hydrate—Caustic Potash** or **Potassa** (U. S.), **Potassa Caustica** (Br.), (KOH), is prepared by the decomposition of potassium carbonate by calcium hydrate (slacked lime).

$$K_2CO_3 + Ca(OH)_2 = CaCO_3 + 2KOH.$$

After these substances have been boiled together, the solution is allowed to settle. It is then poured off, evaporated, and the residue fused in a silver dish (which it does not attack). The fused mass is then poured into moulds. This is called **potash by lime,** and is not pure. To render it purer it is dissolved in alcohol, the solution evaporated to dryness, the residue again melted and cast in silver moulds; this product is **potash by alcohol,** and is free from the chloride and other potassium salts. It is a white, opaque, brittle solid, usually met with in the form of cylindrical sticks, but sometimes in lump. It has a specific gravity of 2.1. It fuses quite easily, and, at high temperatures, volatilizes undecomposed. It is freely soluble in water; less so in alcohol. The solutions give a markedly alkaline reaction, saponify fats, and are strongly caustic. Exposed to the air it absorbs water and carbon dioxide, and is changed into the carbonate. In watery solution it is largely used as a reagent in chemical analysis; it dissolves chlorine, bromine, iodine, sulphur and phosphorus. It decomposes the ammoniacal salts, liberating ammonia; it also decomposes the salts of many of the metals, with the formation of a potassium salt and a hydrate of the metal.

The **haloid salts** of potassium form by direct union of the haloids with the metal, or by saturating the hydrate or carbonate with one of the haloid acids. They all have a bitter, salty taste, are freely soluble in water, and crystallize in cubes. They fuse easily, and are somewhat volatile.

320. **Potassium Chloride** (KCl), occurs native, either pure or mixed with other chlorides. At Stassfurth it is found in large deposits, as **sylvite,** and **carnallite** ($KCl.MgCl_2.CH_2O$). These deposits form the chief source of the potassium compounds.

The chloride crystallizes in anhydrous cubes, of sp. gr. 1.84, closely resembling common salt. 100 parts of water dissolve 30 parts at 0°C. (32°F.), and 0.2738 part more for every degree elevation of temperature.

321. **Potassium Bromide—Potassii Bromidum** (U. S., Br.), (KBr), is generally obtained by dissolving bromine in a solution of potassium hydrate; the bromate also produced in the

reaction is converted into bromide by calcining the product. It has the properties of the other haloid salts, and is used in photography, and in medicine.

322. **Potassium Iodide—Potassii Iodidum** (U. S., Br.), (KI), may be prepared like the preceding, by using iodine instead of bromine. It crystallizes in large, white, translucent cubes, permanent in air, and salty in taste. It dissolves to the extent of 100 parts in 73.5 parts water at ordinary temperatures. Its aqueous solution dissolves iodine in large quantities, forming the **compound tincture** of iodine. It also dissolves many metallic iodides to form double iodides. Its medicinal effects are those of iodine. In chronic poisoning by lead or mercury, it is supposed that it unites with the metals in the blood or tissues, to form soluble iodides, and thus pass them out by the urine.

323. **Potassium Fluoride** (KFl).—Its aqueous solution attacks glass. Is not of much importance to the medical student.

324. **Potassium Cyanide** (KCN), may be obtained either by saturating potassium hydrate with hydrocyanic acid, or by heating potassium ferrocyanide. It is a white, amorphous, deliquescent mass, easily fusible, and smelling of prussic acid. Its solution is very poisonous. Its effects upon the economy are uncertain, but are probably those of hydrocyanic acid. In case of poisoning by it, the stomach should be evacuated and the antidotes of hydrocyanic acid given.

Ferrocyanide and Ferricyanide—see page 179, Art. 293.

325. **Potassium Chlorate—Potassii Chloras** (U. S. P.), ($KCLO_3$).—When a hot, concentrated, solution of potassium hydrate is treated with chlorine gas, the following reaction occurs—

$$6KOH + 3Cl_2 = 5KCl + KClO_3 + 3H_2O.$$

It is usually made by the action of chlorine upon a mixture of calcium hydrate and potassium chloride. By this method a double reaction takes place. Calcium chlorate is first formed.

$$6Ca(OH)_2 + 6Cl_2 = 5CaCl_2 + Ca(ClO_3) + 6H_2O.$$

This then reacts with the potassium chloride.

$$Ca(ClO_3)_2 + 2KCl = 2KClO_3 + CaCl_2.$$

The hot solution is rapidly evaporated, and the residue purified by recrystallization. It crystallizes in shining, transparent plates of the monoclinic system. Soluble in water to the extent of 6.03 parts in 100 at 15.37°C. (59.65°F.), and 24 parts at 104.7°C. (220.5°F.); soluble with difficulty in alcohol. It is cooling and astringent to the taste, fuses at 400°C. (752°F.), giving up a portion of its oxygen and changing to the **perchlorate** ($KClO_4$), which at higher temperatures decomposes into oxygen and potassium chloride. As it gives up oxygen easily, it serves as a valuable oxidizing agent, and as a means of preparing this gas. Mixed with readily oxidizable substances, as carbon, sulphur, phosphorus, sugar, tannin, resins, etc., the mixtures explode when heated or subjected to a sudden shock. The igniting material with which parlor matches are tipped, consists of antimony sulphide and potassium chlorate; rubbed upon a surface coated with red phosphorus, they ignite.

326. **Potassium Hypochlorite** ($KClO$), is formed by the action of chlorine upon a cold solution of potassium hydrate.

$$2KHO + Cl_2 = KCl + KClO + H_2O.$$

It can only be obtained in aqueous solution. If the solution is evaporated, the salt splits up into chloride and chlorate.

$$3KClO = 2KCl + KClO_3.$$

When treated with acids, it yields free chlorine and bleaches strongly. The ordinary solutions used in bleaching (Labarraque's solution and Javelle water), are solutions of impure sodium or potassium hypochlorite.

327. **Potassium Nitrate—Nitre—Saltpetre—Potassii Nitras** (U.S.). (KNO_3), exists native, and is produced artificially whenever nitrogenous organic substances decay in the presence of potassium carbonate. Upon the so-called saltpetre plantations, manures and various animal refuse are arranged in layers with wood ashes and lime, in large heaps, and submitted to the action of the air for two or three years, whereby, from the slow oxidation of the nitrogen, nitrates of potassium and calcium are produced. The heaps are then lixiviated with water and potassium carbonate added to the solution, which contains potassium and calcium nitrates, to convert the last salt into potassium nitrate.

$$Ca(NO_3)_2 + K_2CO_3 = CaCO_3 + 2KNO_3.$$

The calcium carbonate is filtered off, and the solution evaporated. Another method, and probably the one most frequently employed at present, consists in the decomposition of Chili saltpetre, **sodium nitrate,** by means of potassium carbonate, or chloride.

$$NaNO_3 + KCl = NaCl + KNO_3.$$

It crystallizes in large, six-sided, rhombic prisms. 100 parts of water dissolve 244 parts of the salt at 100°C. (212°F.), but at 0° C. (32°F.) only 13 parts. It fuses at 350°C. (662°F.): below a red heat it decomposes into oxygen and potassium nitrite, KNO_2.

The readiness with which it gives up its oxygen when heated in the presence of an oxidizable substance, renders it of value as an oxidizing agent.

Gunpowder is a granular mixture of potassium nitrate, sulphur and charcoal, in such proportion that the nitre contains all the oxygen necessary for the combustion. The following equation expresses approximately the decomposition caused by the burning of powder—

$$2KNO_3 + S + 3C = K_2S + 3CO_2 + N_2.$$

The effect produced, therefore, depends upon the disengagement of carbon dioxide and nitrogen, the volume of which gases is almost 100 times greater than that of the decomposed powder. The heat of the combustion further expands the gases at the time of the explosion.

328. **Potassium Carbonate—Potassii Carbonas** (U. S.) **—Potassæ Carbonas** (Br.)**—Salt of Tartar—Pearlash—** (K_2CO_3),—exists in mineral waters, in the animal economy, and as the principal ingredient of wood ashes. The plants absorb potassium salts from the earth and convert them into salts of the organic acids. When the plants are burned, the organic acids are destroyed and potassium carbonate produced, which is obtained by lixiviation of the ashes, and evaporation. This method is not much employed at present. The immense deposits in Stassfurth and Gallicia afford an almost inexhaustible supply of potassium salts. It occurs commercially as a white, granular, deliquescent powder, freely soluble in water, the solution having a caustic taste and an alkaline reaction.

329. **Potassium Bicarbonate—Hydropotassic Carbonate—Potassii Bicarbonas** (U. S.)**—Potassæ Bicarbonas** (Br.), ($KHCO_3$).—When carbon dioxide is passed through a concentrated solution of potassium carbonate, it is absorbed and potassium bicarbonate produced.

$$K_2CO_3 + H_2O + CO_2 = 2KHCO_3.$$

This salt crystallizes in oblique, rhombic prisms, of the monoclinic system. It dissolves in 3 to 4 parts of water; the solution is faintly alkaline but not caustic. The substance that is still extensively used in some parts of the country in baking, under the name **saleratus**, is this, or the corresponding sodium salt. It "raises" the bread by the action of heat in setting free the carbon dioxide, and leaving potassium (or sodium) carbonate, which, by its strongly alkaline reaction, may cause digestive disturbances.

330. **Sulphides.**—Five are known: K_2S, K_2S_2, K_2S_3, K_2S_4, and K_2S_5; also a sulphydrate, KSH. The latter is prepared by the action of hydrogen sulphide upon potassium hydrate.

$$KOH + H_2S = KSH + H_2O.$$

The **Pentasulphide** (K_2S_5), **Liver of Sulphur—Potassii Sulphuretum** (U. S.)—**Potassa Sulphurata** (Br.), is obtained by fusing potassium carbonate with an excess of sulphur. It decomposes readily, and in contact with hydrochloric acid gives off hydrogen sulphide.

331. **Potassium Sulphate—Dipotassium Sulphate—Potassii Sulphas** (U. S.)—**Potassæ Sulphas** (Br.), (K_2SO_4), is found in the Stassfurth mines, in plant ashes, and in solution in mineral waters. It is obtained by the action of sulphuric acid upon potassium chloride, as a by-product in some chemical manufacturing processes. It crystallizes without water, in small, rhombic prisms, of a bitter, salty taste, and is soluble in 10 parts of water at ordinary temperatures.

332. **Hydropotassium Sulphate—Monopotassic Sulphate—Acid Sulphate** ($KHSO_4$), is formed as a by-product in the manufacture of nitric acid from potassium nitrate; crystallizes in large, rhombic tables, and is very readily soluble in water. At about 200°C. (392°F.) it fuses, loses water, and is converted into the pyrosulphate ($K_2S_2O_7$).

333. **Sulphites.**—Three are known: K_2SO_3, $KHSO_3$ and $K_2S_2O_5$. **Potassium Sulphite—Neutral Potassium Sulphite—Potassium Sulphis** (U. S.), (K_2SO_3).—This salt crystallizes in oblique rhombic octahedra, which dissolve readily in water, and have a sulphurous odor. When in solution, if exposed to the air, it absorbs oxygen, and is converted into the sulphate.

334. **Potassium Acetate—Potassii Acetas** (U. S.)—**Potassæ Acetas** (Br.), ($KC_2H_3O_2$), exists in the juices of plants. It is obtained by neutralizing acetic acid with potassium carbonate or bicarbonate. It crystallizes in shining needles, deliquescent, and very soluble in water.

335. **Oxalates.**—Three are known to exist: **Potassium Oxalate—Neutral Oxalate** ($K_2C_2O_4 + Aq.$), formed by saturating oxalic acid with potassium carbonate. **Hydropotassium Oxalate—Monopotassium Oxalate—Binoxalate of Potash** (KHC_2O_4). **Potassium Quadroxalate** ($KHC_2O_4.\ C_2O_4H_2 + 2Aq.$). A mixture of these two salts is known as **salt of lemon**, or **salt of sorrel**, and is used for bleaching straw and to remove ink stains. In appearance it closely resembles Epsom salt, and has caused many cases of oxalic acid poisoning, being taken by mistake for that salt.

336. **Tartrates.—Potassium Tartrate—Soluble Tartar—Neutral Tartrate of Potash—Potassii Tartras** (U. S.)—**Potassæ Tartras** (Br.), ($K_2C_4H_4O_6$), is a white, crystalline powder, very soluble

in water; soluble in 240 parts alcohol. **Hydropotassium Tartrate—Cream of Tartar—Potassii Bitartras** (U.S.), **Potassæ Bitartras** (Br.), ($KHC_4H_4O_6$).—A brown-red, crystalline crust is obtained from the bottom and sides of wine-casks after fermentation has taken place; this is known in commerce as **argol**, or **crude tartar**, and is composed in great part of potassium bitartrate, with tartrate of lime and coloring matter.

The argol is boiled with water, or heated in a closed digester by superheated steam. The latter process renders the calcium tartrate insoluble and separates it almost completely from the cream of tartar, which goes into solution. The solution thus obtained is allowed to cool and crystallize; the crystals are redissolved in hot water, treated with animal charcoal, to remove coloring matters, filtered, and again crystallized. The product of this process is almost chemically pure acid potassium tartrate.

It crystallizes in hard, opaque, rhombic prisms, very sparingly soluble in water, still less so in alcohol. Its solution is acid to the taste, and to litmus paper. It is largely used in baking, combined with sodium bicarbonate, the two substances reacting upon each other to form Rochelle salt, with liberation of carbon dioxide. **Baking powders** are extensively used at present, instead of yeast, for bread making. In all of them the action depends upon the decomposition of sodium bicarbonate by some salt having an acid reaction, or by a weak acid.

In addition to the bicarbonate, and the starch added to preserve them, many of them contain either tartaric acid, alum, or acid phosphates of calcium, instead of cream of tartar.

Some of the reactions that take place to set free the carbon dioxide are the following:—

$$1.\ \underset{\substack{\text{Potassium}\\\text{bitartrate.}}}{KHC_4H_4O_6} + \underset{\substack{\text{Sodium}\\\text{bicarbonate.}}}{NaHCO_3} = \underset{\substack{\text{Sodio-potassium}\\\text{tartrate.}}}{NaKC_4H_4O_6} + \underset{\text{Water}}{H_2O} + \underset{\substack{\text{Carbon}\\\text{dioxide.}}}{CO_2}.$$

$$2.\ \underset{\substack{\text{Aluminium potassium}\\\text{alum.}}}{Al_2(SO_4).K_2SO_4} + \underset{\substack{\text{Sodium}\\\text{bicarbonate.}}}{6NaHCO_3} = \underset{\substack{\text{Potassium}\\\text{sulphate.}}}{K_2SO_4} + \underset{\substack{\text{Sodium}\\\text{sulphate.}}}{3Na_2SO_4} +$$
$$\underset{\substack{\text{Aluminium}\\\text{hydrate.}}}{Al_2H_6O_6} + \underset{\substack{\text{Carbon}\\\text{dioxide.}}}{6CO_2}.$$

$$3.\ \underset{\substack{\text{Calcium}\\\text{biphosphate.}}}{CaH_4(PO_4)_2} + \underset{\substack{\text{Sodium}\\\text{bicarbonate.}}}{2NaHCO_3} = \underset{\substack{\text{Hydro-calcium}\\\text{phosphate.}}}{CaHPO_4} + \underset{\substack{\text{Sodium}\\\text{phosphate.}}}{NaHPO_4} +$$
$$\underset{\text{Water.}}{2H_2O} + \underset{\substack{\text{Carbon}\\\text{dioxide.}}}{2CO_2}.$$

An artificial powder may be made by intimately mixing two parts cream of tartar with one of sodium bicarbonate, and adding a little flour or starch.

337. **Sodium Potassium Tartrate—Rochelle Salt—Potassii et Sodi Tartras** (U. S.), ($NaKC_4H_4O_6 + 4Aq.$), is prepared by boiling acid potassium tartrate with sodium carbonate. It forms large, transparent, prismatic, slightly efflorescent

crystals, soluble in 2½ parts of cold water, saline, and slightly bitter to the taste, and neutral in reaction.

338. **Potassium Antimonyl Tartrate — Tartarated Antimony—Tartar Emetic—Antimonii et Potassi Tartras** (U. S.), ($SbOKC_4H_4O_6$), is prepared by boiling a solution of cream of tartar with antimonious oxide. Its crystals are transparent, right rhombic octahedra, efflorescing in the air. It is quite soluble in water, the solution having a nauseating, metallic taste. It is poisonous, and has even caused death when applied to the skin as a local irritant and vesicant.

RUBIDIUM and CÆSIUM.

Rb = 85.4. Cs = 132.6.

339. These rare metals were discovered in 1860 by Bunsen and Kirchoff, by means of the spectroscope. Both elements were named from the color of their lines in the spectrum (**rubidius**, dark red, and **cæsius**, sky blue). They occur in small quantities, widely distributed, often accompanying potassium. With platinum chloride they form double chlorides ($PtCl_4.2RbCl$). Their compounds have found no use in medicine, as yet.

AMMONIUM COMPOUNDS.

340. **Ammonium** (NH_4).—This radical has only a hypothetical existence, never having been isolated. But there are many reasons to believe that it does actually exist in combination in the ammonium compounds, and that in these compounds it plays the rôle of a metal resembling sodium and potassium. The oxide of this radical has not been separated.

341. **Ammonium Hydrate** (NH_4OH) is believed to exist in solution in the ordinary aqua ammoniæ, although, when the attempt is made to isolate it, decomposition ensues.

Aqua Ammoniæ Fortior (U. S.) has a sp. gr. of 0.900, and contains 249.5 grams NH_3 per litre.

Aqua Ammoniæ (U. S.) is of sp. gr. 0.960, and contains 93.1 grams NH_3 per litre.

342. **Ammonium Chloride—Ammonium Muriate—Sal-Ammoniac—Ammonii Chloridum** (U. S., Br.), (NH_4Cl), was formerly obtained by the dry distillation of camels' dung, and is sometimes found in volcanic regions. At present it is prepared chiefly by saturating the ammonia water from gas works with hydrochloric acid, evaporating the solution to dryness, and subliming the residue in iron vessels. Prepared in this way, it is

a compact, tough, fibrous mass, which dissolves in 2.7 parts cold, and one part boiling water. It crystallizes from its solution in small octahedra or cubes, of a sharp, salty taste, and neutral reaction. When heated, it volatilizes without fusing; at the same time a dissociation into NH_3 and HCl occurs, but on cooling, these products recombine into ammonium chloride.

This salt exists in minute quantities in the gastric juice of various animals. The urine, saliva, and tears also contain some ammonium compound, which is said to be the chloride.

343. **Ammonium Bromide — Ammonii Bromidum** (U. S., Br.), ((NH_4) Br.), may be prepared by direct combination of ammonia and hydrobromic acid, or by decomposing ferrous bromide with aqua ammoniæ. It forms a white, granular powder, or large prisms, which turn yellow on exposure to air, and possess a saline, pungent taste, and neutral reaction. It dissolves in 1.5 parts of water, and volatilizes without decomposition.

344. **Ammonium Iodide—Ammonii Iodidum** (U. S.), ($(NH_4)I$), is prepared by the action of hydriodic acid upon ammonia, or by the double decomposition of potassium iodide and ammonium sulphate. It forms cubic crystals, which are deliquescent, and is soluble in 0.60 parts of water. It decomposes in air, turning yellow and emitting the odor of iodine.

345. **Ammonium Sulphide** ($(NH_4)S_2$) is a white, crystalline solid, formed by mixing dry hydric sulphide and ammonia at a low temperature (—18°C. (about 0°F.). It is usually prepared and used as a yellow solution, by mixing the sulphydrate with ammonium hydrate. It dissolves sulphur and the sulphides of arsenic, tin and antimony, and is used in analysis for this purpose.

Ammonium Sulphydrate (NH_4SH) is prepared by saturating a solution of ammonium hydrate with hydrogen sulphide (sulphuretted hydrogen). It is colorless at first, but becomes yellow, from decomposition, on exposure. It is used in laboratories as a reagent. Acids decompose both these sulphides, setting free sulphur.

346. **Ammonium Carbonates.**—Three are known: **Ammonium Carbonate—Neutral Ammonium Carbonate** ($(NH_4)_2CO_3$), may be prepared as a crystalline powder, by passing ammonia gas through a concentrated solution of the sesquicarbonate. Exposed to the air it splits up into ammonia and the acid carbonate, NH_4HCO_3.

Hydro-ammonium Carbonate—Acid Ammonium Carbonate (NH_4HCO_3), is obtained when a solution of ammonium hydrate or sesquicarbonate is saturated with carbon dioxide. It forms large, rhombic crystals, which are quite soluble in water. At 60°C. (140°F.) it is decomposed into ammonia and carbon dioxide.

Ammonium Sesquicarbonate—Sal-Volatile—Ammonii Carbonas (U. S.)—**Ammoniæ Carbonas** (Br.), $(NH_4)_4H_2)(CO_3)_3$, is the commercial carbonate of ammonia, and was formerly prepared by the dry distillation of bones, horns, and other animal substances. It is at present prepared by heating a mixture of ammonium chloride or sulphate with calcium carbonate, and condensing the volatilized product. So prepared, it sublimes as a white, transparent, hard mass, having an ammoniacal odor and an alkaline reaction. On exposure to the air, it gives off ammonia and carbon dioxide. The carbonates of ammonia are very unstable.

347. **Ammonium Nitrate—Ammonii Nitras** (U. S.), $(NH_4)NO_3)$, is prepared by neutralizing nitric acid with ammonium hydrate or carbonate. It crystallizes in flexible, six-sided prisms, without water; dissolves in 0.5 parts water at 18°C. (64.4°F.), and fuses at 150°C. (302°F.). When heated to 210°C. (410°F.), it decomposes, with the formation of nitrous oxide, or laughing gas, and water.

$$(NH_4)NO_3 = N_2O + 2H_2O.$$

348. **Ammonium Sulphate—Neutral Ammonium Sulphate—Ammonii Sulphas** (U. S.), $(NH_4)_2SO_4)$, may be obtained by saturating the ammonia water from gas works with sulphuric acid. It forms rhombic crystals, soluble in two parts of cold, and one part of hot water. At 140°C. (284°F.) it fuses, and at higher temperatures it decomposes into ammonia, nitrogen, water and ammonium sulphite.

349. **Ammonium Acetate** $(NH_4)C_2H_3O_2)$ is formed when acetic acid is saturated with ammonia or ammonium carbonate. It is seldom seen except in solution in water. The aqueous solution is used in medicine, as the **Liq. Ammonii Acetatis,** or **Spirit of Mindererus.**

Other salts in use are the **benzoate, phosphate and valerianate,** all white crystalline salts.

350. **Action on the Economy.**—In large quantities, or by prolonged use, ammonia and its salts are poisonous. Ammonia, if inhaled, acts as a severe irritant upon the air passages, causing dyspnœa, pain, suffocation, and even death. The treatment in cases of poisoning, consists in neutralizing the alkali by dilute

acids, or the vapor of acetic or dilute hydrochloric acids may be inhaled. Two drachms of a strong solution of ammonium hydrate have proved fatal.

METALS OF THE ALKALINE EARTHS.

Calcium = 40. **Strontium** = 87.5. **Barium** = 137.2.

351. The members of this group are called metals of the alkaline earths because their oxides resemble, on the one hand, the oxides of the alkalies, and, on the other, the real earths (aluminia, etc.). Like the potassium group, their properties and chemical energy increase gradually with their atomic weights. Their basic properties also become greater with their atomic weights. Thus, barium decomposes water more energetically, and oxidizes more readily, than strontium or calcium. Barium hydrate is likewise the strongest base. It is quite soluble in water. It fuses without decomposition upon ignition, and absorbs carbon dioxide rapidly from the air. Calcium hydrate possesses weaker basic properties, is difficultly soluble in water, and, when ignited, breaks up into water and calcium oxide. Strontium stands between barium and calcium. Thus, although these metals resemble the alkaline metals, both in their free state and as hydrates, they essentially differ from them in the insolubility of their carbonates and phosphates, and still more their sulphates. Barium sulphate is insoluble in water and acids.

CALCIUM.

Ca = 40.

352. **Occurrence.**—This metal forms one of a class of elements most widely distributed in nature. Its carbonate (limestone, marble and chalk), its sulphate (gypsum and alabaster), and its phosphate, fluoride, and silicate, are common minerals.

Preparation.—Calcium may be obtained from the fused chloride, by electrolysis; or, by heating calcium iodide with sodium, or calcium chloride with sodium and zinc.

Properties.—A light, shining, yellow, ductile metal. It fuses at a red heat, does not sensibly volatilize, and in the air burns with a brilliant reddish-yellow light. It does not undergo decomposition in dry air, but in moist air covers itself with a layer of hydrate. Its specific gravity is 1.984.

353. **Calcium Oxide—Lime—Calx** (U. S., Br.), (CaO.), is obtained pure by igniting the carbonate or nitrate. On a large scale, it is prepared commercially by burning the natural carbonate (limestone or marble) in rude stone furnaces, called lime

kilns. It is a grayish-white, amorphous solid, alkaline and caustic; sp. gr. 2.3; it does not fuse at any temperature at our command. When the oxy-hydrogen flame is thrown upon it, it becomes incandescent, and emits an extremely intense white light. (Calcium, or Drummond light.) It combines energetically with water to form the hydrate, the process known as **slacking**, and is attended with the evolution of much heat. Exposed to the air it attracts moisture and becomes **air-slacked.**

354. **Calcium Hydrate—Slacked Lime—Calcis Hydras** (Br.), ($Ca(OH)_2$, is a dry, white powder, odorless, and alkaline in reaction. It is slightly soluble in cold and less so in hot water. Exposed to the air it absorbs carbon dioxide and forms carbonate. Ordinary mortar is a mixture of slacked lime, water and quartz sand. The hardening of mortar depends upon three causes: 1st. The natural evaporation of the water. 2d. The absorption of carbon dioxide from the air, and the formation of calcium carbonate. 3d. The action of the hydrate upon the silicic acid of the sand, producing calcium silicate. This last reaction takes place slowly; hence the hardness of old mortars. Hydraulic mortar, or **cement,** contains calcium oxide, aluminium silicate and quartz powder; its hardening depends principally upon the formation of calcium and aluminium silicates.

355. **Lime Water—Liquor Calcis** (U. S., Br.), is a clear solution of the calcium hydrate in water. Cane-sugar increases the solubility of calcium hydrate in water, with which it forms a saccharate. The British **Liquor calcis saccharatus** is a solution of calcium hydrate in a strong solution of cane-sugar. When lime-water contains an excess of the hydrate, rendering it turbid, it is called milk of lime.

356. **Calcium Chloride—Calcii Chloridum** (U. S., Br.), ($CaCl_2$), is prepared by the action of hydrochloric acid upon marble. It crystallizes with 6 molecules of H_2O, in large, six-sided prisms, which are bitter, deliquescent, and very soluble in water. If heated, it melts in its water of crystallization with some loss of water. Above 200°C. (392°F), it becomes anhydrous. The dry salt is a white, porous mass, which fuses at a red heat and solidifies to a crystalline mass, which rapidly absorbs water, and is used as a drying agent for gases and liquids. **Calcium iodide** and **bromide** are very similar to the chloride.

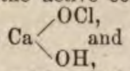
357. **Chloride of Lime—Bleaching Powder—Calx Chlorinata** (U. S.), **Calx Chlorata** (Br.), is a mixture of calcium chloride ($CaCl_2$), and calcium hypochlorite ($Ca(ClO)_2$),

with some water. The hypochlorite is the active principle. It is prepared by passing chlorine over slacked lime.

From the analogous action of chlorine upon sodium or potassium hydrate, we may express the reaction in the case of calcium hydrate by the following equation:—

$$2Ca(OH)_2 + 2Cl_2 = Ca(OCl)_2 + CaCl_2 + 2H_2O.$$

According to this equation, the completely chlorinated chloride of lime must contain 48.9 per cent of chlorine, which is never the case, as a portion of the calcium hydrate always remains unchanged. The exact constitution of chloride of lime is in doubt; but from more recent observations, it is believed by some that the active constituent of chloride of lime is a basic calcium hypochlorite $Ca\begin{cases}OCl,\\ OH,\end{cases}$ and the following reaction takes place when chlorine acts upon calcium hydrate:—



$$3Ca(OH)_2 + 2Cl_2 = 2CaO_2HCl + CaCl_2 + 2H_2O.$$

Calculating from this equation, completely saturated chloride of lime only contains 39 per cent. chlorine, which is found to be actually the case.

Chloride of Lime is a grayish-white, porous powder, having a bitter, acrid taste, and a chlorine-like odor; it is alkaline in reaction, soluble in cold and decomposed by boiling water. It slowly decomposes in the air, the carbon dioxide liberating hypochlorous acid; this decomposition is hastened by sunlight and heat. Dilute mineral acids decompose it very rapidly, with liberation of chlorine. The application of chloride of lime for bleaching and disinfecting purposes, depends upon this production of free chlorine. The amount that will be set free by acids, is called the **active chlorine,** which in good chloride of lime should be at least 25 per cent.

358. **Calcium Sulphate** ($CaSO_4$) occurs very abundantly in nature as **gypsum**, in right rhombic prisms, combined with two molecules of water. The anhydrous salt forms the mineral **anhydrite.** It is very sparingly soluble in water. One part dissolves at ordinary temperatures in 400 parts H_2O. Ground gypsum is used in the arts under the name of **terra alba.** When heated to 200°C. (392°F.), it parts with its water, becoming converted into an opaque white mass, which, when ground, is called **Plaster-of-Paris.** This powder, mixed with water, takes up two molecules and hardens into a stone-like solid. Upon this property depends the usefulness of plaster for making moulds, figures, and immovable surgical dressings.

359. **Tricalcium Phosphate—Tribasic,** or **Neutral,** or **Bone Phosphate—Calcis Phosphas Precipitata** (U. S.), **Calcis Phosphas** (Br.), ($Ca_3(PO_4)_2$), is found in rocks and

soils, in guano, in all plants, and in every tissue and fluid of animal bodies.

It may be prepared pure by dissolving bone ash in hydrochloric acid, filtering, and precipitating with ammonium hydrate; or, by a double decomposition between calcium chloride and an alkaline phosphate. It is a gelatinous mass when first precipitated; but, after drying, a white, amorphous powder. It is insoluble in water, but readily soluble in dilute acids, even acetic; also in water charged with carbon dioxide. An impure tricalcium phosphate, prepared by burning bones, is known as **bone-ash.**

360. **Dicalcium Phosphate** ($Ca_2HPO_4 + 2H_2O$) separates as an amorphous insoluble precipitate, when disodium phosphate is added to a solution of calcium chloride mixed with some acetic acid.

361. **Monocalcium Phosphate—Acid Calcium Phosphate—Superphosphate of Lime** ($Ca(H_2PO_4)_2$), is found in brain tissue and in acid animal fluids. It is produced by the action of sulphuric or hydrochloric acid upon the first two phosphates, and is manufactured as a fertilizer, mixed with calcium sulphate, by decomposing bones with sulphuric acid. At a temperature of 200°C. (392°F.), it splits up into pyrophosphate, metaphosphoric acid and water.

$$2Ca\,(H_2PO_4)_2 = Ca_2P_2O_7 + 2HPO_3 + 3H_2O.$$

When this mixture is ignited with charcoal, the metaphosphoric acid is reduced to phosphorus.

362. **Calcium Carbonate—Calcis Carbonas** (U. S., Br.), ($CaCO_3$), is of exceedingly wide distribution in nature. It exists sometimes in enormous deposits, as limestone, marble, chalk, Iceland spar, and as the mineral basis of the corals, shells of the crustacea, mollusks, etc. Chalk is a comparatively pure, amorphous calcium carbonate, made up of microscopic shells. The precipitated chalk of the Pharmacopœia is prepared by precipitating calcium chloride with sodium carbonate, filtering off and washing with water. **Prepared chalk—creta preparata** (U. S., Br.), is a native chalk, purified by **elutriation,** which consists in grinding the chalk in water, allowing the mixture to subside, decanting the upper portion, and collecting and drying the finer particles.

Calcium carbonate is nearly insoluble in pure water, but dissolves readily in water containing carbon dioxide; hence, we find it dissolved in nearly all natural waters, as an acid or bicarbonate ($CaH_2(CO_3)_2$), giving rise to temporary hardness. Boiling, agitation, or free exposure to the air may decompose this salt, and deposit the ordinary calcium carbonate. Upon this depends the formation of lime scales, stalactites, boiler incrustations, and similar deposits.

363. **Calcium Oxalate** (CaC_2O_4) is found in the juice of some plants and in the urine. It may be obtained, as a fine, white, crystalline powder, by adding any soluble oxalate to a soluble calcium salt in neutral or alkaline solution. It is insoluble in water and acetic acid, but soluble in the mineral acids. In many diseased conditions which produce deficient oxidation or excessive production of acids, (lung disease and acid dyspepsia), it occurs in considerable quantities in the urine, and gives rise to **oxaluria**, or the oxalic acid diathesis. This salt frequently forms calculi, which present irregular projections and have received the name of mulberry calculi. Excessive saccharine diet, or excessive consumption of certain vegetables, as tomatoes, rhubarb, etc., increases the production of calcium oxalate.

364. **Physiological Effects and Uses.**—The calcium salts play an important part in the animal economy. The phosphates are found in every tissue and fluid of the body, but most abundant in the bones and teeth; the former containing from 55 to 59 per cent., and the latter, including the carbonate, 72 per cent. As the salts of lime are insoluble in alkaline fluids, various theories have been put forward to explain its state in the blood and other alkaline fluids. It seems certain that the calcium of blood serum does not exist as phosphate, but as some soluble salt, or albumin compound, soluble in alkaline fluids. The calcium phosphate of the urine remains in solution as long as that fluid is acid, but separates as an amorphous or crystalline sediment as soon as it undergoes alkaline fermentation. Alkaline urine is always turbid. When taken internally, the calcium salts produce effects similar to those of sodium and potassium, but milder. They have a mild astringent effect.

STRONTIUM.

$Sr = 87.5.$

365. Strontium is rather sparingly found in nature. The principal minerals are strontianite ($SrCO_3$) and celestite ($SrSO_4$). It is a brass-yellow, lustrous metal, which resembles calcium in its properties, as also do its compounds. This element exhibits most of the properties of calcium and barium. Its compounds impart a red tinge to the flame, and for this reason the nitrate is used as a constituent of red fires. The solubility of its

sulphate stands between that of the calcium and barium sulphates. It has been used in medicine as an alterative. Its salts are poisonous only in large quantities.

BARIUM.

$Ba = 137.2.$

The element itself is not of interest to the medical student.

366. **Barium Oxide—Baryta** (BaO), is obtained by calcining the nitrate. It is a grayish-white, caustic powder, fusible in the oxyhydrogen flame. It unites with water, with the evolution of much heat, to form the hydrate BaH_2O_2; this dissolved in water forms baryta water.

There is a **barium peroxide** known (BaO_2), which is decomposed by dilute acids, with the production of hydrogen peroxide.

367. **Barium Chloride—Barii Chloridum** (U. S.), ($BaCl_2 + 2Aq.$), is prepared by the action of hydrochloric acid upon the native sulphide or carbonate. It is used as a reagent for soluble sulphates, giving a white precipitate, insoluble in acids or water.

It has been used in medicine as an alterative and anthelmintic.

368. **Barium Nitrate** ($BaNO_3$) forms anhydrous octahedra of the regular system, soluble in water, and used as a constituent of pale green theatre fires.

369. **Barium Sulphate** ($BaSO_4$), **Heavy Spar,** or **Barite,** occurs in nature in rhombic prisms and amorphous; sp. gr. 4.6. It is obtained by the action of sulphuric acid upon barium salts, as a white amorphous powder, nearly insoluble in acids and water. It is used as an adulterant of white paint, Paris green, and a variety of other commercial products.

370. **Barium Carbonate** ($BaCO_3$) also occurs native as witherite. It precipitates from solutions of barium salts, as a white, amorphous powder, when they are acted upon by soluble carbonates.

371. **Physiological Effect of Barium Salts.** All the soluble compounds of barium, as well as those that are converted into soluble compounds in the stomach, are poisonous. Whenever a poisonous dose has been taken, the patient should take some soluble sulphate (as Epsom or Glauber's salt), followed by an emetic. The symptoms of poisoning are pain in stomach, prostration, dilated pupils, loss of voice, sight or hearing, excessive micturition, and other very prominent nervous symptoms. Post-mortem, inflammation and great friability of stomach, in most cases, and invariably great inflammation of the rectum are found.

METALS OF THE MAGNESIUM GROUP.

Magnesium = 24. **Zinc** = 65.2. **Cadmium** = 112.

372. These metals form but one oxide, with a corresponding basic hydrate, and a series of salts in which the atoms possess a quantivalence of two. According to some authorities, the rare element **beryllium** is included in the group.

MAGNESIUM.

$Mg = 24.$

373. **Occurrence.** A metal, found abundantly in nature, usually accompanying calcium.

Dolomite, an isomorphous mixture of calcium and magnesium carbonates, forms the so-called magnesium limestone. Asbestos, serpentine, meerschaum, talc or soapstone, and hornblende are native silicates. Nearly all natural waters contain some of the soluble salts of magnesium, which impart hardness to them.

374. **Preparation.** It may be obtained either by electrolysis of the chloride, or by heating the same compound with sodium.

In the arts it is prepared by fusing the double chloride of sodium and magnesium with metallic sodium.

$$MgCl_2NaCl + 2Na = 3NaCl + Mg.$$

It is then purified by distillation in an atmosphere of hydrogen.

375. **Properties.**—Magnesium is a brilliant white metal, of a sp. gr. of 1.75; very tenacious and ductile. It fuses at a dull red heat, and at a bright red heat it distills. It oxidizes but slightly in the air, at ordinary temperatures, but when heated, it burns with an intensely brilliant bluish-white light, owing to the incandescence of the non-volatile magnesium oxide. The flame of burning magnesium is rich in chemically active or actinic rays; hence, it is much employed for photographing in dark caves and subterraneous chambers. It combines directly with chlorine, sulphur, phosphorus, arsenic and nitrogen. Soluble in dilute acids, but not in alkalies. It is slowly oxidized by boiling water.

376. **Magnesium Oxide—Calcined Magnesia—Magnesia** (U. S., Br.), (MgO), is formed by the combustion of the metal, or by the ignition of the carbonate, hydrate, or nitrate. It is a very light, white powder, without odor or taste, and has a feeble alkaline reaction. It is soluble in dilute acids.

A compact variety, prepared by heating the nitrate or chloride to bright redness, and no higher, exhibits remarkable hydraulic properties. If moistened with water to a paste, it quickly hardens to a compact white solid, of great hardness and durability. If it be mixed with an equal part of marble dust or chalk and moistened, it may be moulded into any desired shape, and on being placed in water, it "sets" into an extremely hard mass. It has been used as a filling for decayed teeth.

377. **Magnesium Hydrate — Hydrated Magnesia** $Mg(HO)_2$, is formed from any soluble magnesium salt by precipitating with sodium or potassium hydrate. It is almost insoluble in water and alkalies, but soluble in ammonium salts with the formation of double salts. A mixture holding it in suspension in water, known as **milk of magnesia,** is used in medicine as a laxative and an antidote for acid poisons.

378. **Magnesium Chloride** ($MgCl_2$) exists in small quantities in many mineral springs. It may be obtained by dissolving the carbonate or oxide in hydrochloric acid. It forms deliquescent crystals with 6 molecules of H_2O, isomorphous with calcium chloride. The anhydrous chloride is one of the most deliquescent substances known.

379. **Magnesium Sulphate — Epsom Salt — Seidlitz Salt—Magnesii Sulphas** (U. S.)—**Magnesiæ Sulphas** (Br.), ($MgSO_4$), is found in solution in sea waters and in many mineral springs, especially those belonging to the class of **bitter waters.**

It is prepared by the action of sulphuric acid upon magnesium carbonate. At ordinary temperatures it crystallizes with 7 molecules of H_2O, in four-sided, rhombic prisms, very readily soluble in water. When heated, it fuses and parts with its water of crystallization up to 150°C. (302°F.), when it has lost all but one molecule; this it finally parts with when heated to 200°C. (392°F.) **One** molecule of water, therefore, is more closely combined than the rest. This is known as the **Water of Constitution.**

380. **Magnesium Phosphates.**—These resemble the calcium phosphates, which they generally accompany in the animal body, though usually existing in smaller quantity. Magnesium also forms double phosphates, one of which, the **Ammonio-magnesium phosphate,** or **triple-phosphate** ($MgNH_4PO_4 + 6H_2O$), is precipitated when an excess of an alkaline phosphate and of ammonia is added to a solution containing magnesium. When the urine becomes ammoniacal from the decomposition of urea, this salt is precipitated, as urine always contains alkaline phosphates and magnesium salts. Being practically insoluble, especially in the presence of excess of phosphates

and of ammonia, it is usually deposited from the urine as a sediment, in the shape of modified right rhombic prisms, which, under the microscope, resemble the shape of a coffin lid. This sometimes takes place in the bladder, and if some body is present that will act as a nucleus, the so-called **fusible calculus** may form.

381. **Magnesium Carbonate — Neutral Carbonate** ($MgCO_3$), occurs in nature as **magnesite**, and, combined with calcium carbonate, as **dolomite.** On adding an alkaline carbonate to an aqueous solution of a magnesium salt, magnesium carbonate is not produced, as most other carbonates would be under similar circumstances, but some carbon dioxide escapes, and a white precipitate falls, which is a mixture of magnesium carbonate and hydrate, or **magnesia alba.**

382. **Tetramagnesium Carbonate—Magnesia Alba—Magnesii Carbonas** (U. S.), **Magnesiæ Carbonas** (Br.), ($3(MgCO_3).MgH_2O_2 + 3H_2O$), occurs in commerce in light white cubes, composed of an amorphous or partly crystalline powder. It is prepared by precipitating a solution of magnesium sulphate with one of sodium carbonate. A hot concentrated solution should be used, and the liquid boiled after precipitation. This compound varies in constitution according to the length of time that the boiling has continued, and the presence or absence of excess of sodium carbonate. It is very slightly soluble in water, but quite soluble in solutions of ammonium chloride.

ZINC.

$Zn = 65.2.$

383. **Occurrence.** The native compounds of the heavy metals are termed **ores.** The most common zinc ores are **Smithsonite** ($ZnCO_3$.), **Calamine,** (**Silicate**) and **Sphalerite** or **Blende** (ZnS). These, like most ores, have a high specific gravity, and usually are found in veins in the older crystalline rocks.

384. **Preparation.** To obtain the metal, one of its ores, usually the carbonate or sulphide, is converted into an oxide by roasting, heating to a high temperature in the air. This oxide is then mixed with carbon and ignited in cylindrical earthenware tubes, reducing the oxide thus:

$$ZnO + C = Zn + CO.$$

The free zinc is then distilled off.

385. **Properties.** A bluish-white metal, roughly crystalline or granular; its specific gravity is 6.862, if cast; 7.215, if rolled.

It is brittle at ordinary temperatures and can be pulverized; at 100°C. (212°F.) to 150°C. (302°F.), it is malleable and ductile, and may be rolled into thin sheets. At 200°C. (392°F.), it again becomes brittle. It fuses at 412°C. (773°F.), and distills at 1040°C. (1872°F.). In moist air it becomes coated with a thin layer of basic carbonate. When heated in the air, it burns with a very intense bluish-white light, with the formation of oxide. It dissolves readily in dilute acids with evolution of hydrogen. Concentrated sulphuric acid does not dissolve zinc. It is soluble in sodium, potassium and ammonium hydrates. On account of the slight action of the air upon it, zinc meets with extensive application in architecture, and for galvanizing or coating iron.

386. **Zinc Oxide—Zinci Oxidum** (U. S., Br.), (ZnO), may be prepared by igniting the precipitated basic carbonate, or by burning the metal in a current of air. When obtained in the former way, it forms a soft, white, tasteless and odorless powder. When produced by burning the metal, it occurs as a white, voluminous, flocculent mass, formerly called **flores zinci** or **lana philosophica.** It neither fuses, volatilizes, nor decomposes by heat, and is insoluble in neutral solvents. It is used in the arts as a white pigment, and is not darkened by hydrogen sulphide, as white lead is.

387. **Zinc Hydrate** ($Zn(OH)_2$) is formed, as a white amorphous powder, by precipitating an aqueous solution of a zinc salt, by alkalies. It is soluble in alkaline hydrates and in solutions of ammonium salts. When heated, it decomposes into zinc oxide and water.

388. **Zinc Chloride—Butter of Zinc—Zinci Chloridum** (U. S., Br.), ($ZnCl_2$), is obtained by heating zinc in a stream of chlorine, by dissolving zinc in hydrochloric acid, and by the distillation of zinc sulphate with calcium chloride. It forms a soft white mass, which is very deliquescent; is fusible and volatile. It is extremely soluble in water, freely so in alcohol. The solution has a burning metallic taste, destroying animal and vegetable tissues, and possessing strong dehydrating properties. It forms a series of double salts. The double chloride of zinc and ammonium is sometimes used in soldering, to dissolve metallic oxides from the surface of metals to be soldered.

389. **Zinc Sulphate—White Vitriol—Zinci Sulphas** (U. S., Br.), ($ZnSO_4$), is formed by dissolving zinc, or its oxide, sulphide, or carbonate, in sulphuric acid. At temperatures below 30°C. (86°F.), it crystallizes with 7 Aq.; at 30°C, (86°F), with 6 Aq.; between 40°C. (104°F) and 50° (122°F), with 5

Aq. The most common salt is that with 7 Aq., which occurs in rhombic crystals resembling magnesium sulphate; and is freely soluble in water. It is used in medicine as an emetic and astringent.

390. **Toxicology.** The compounds of zinc that are soluble in the digestive fluids are all irritant poisons. Solutions of the chloride (used by tinsmiths, in embalming, and as a disinfectant in Burnett's fluid,) are also very corrosive. The antidotes are alkaline carbonates, soap, albumen and mucilage. Solutions containing sodium chloride or organic acids act as solvents upon metallic zinc; consequently, symptoms of poisoning, more or less marked, are apt to follow eating of acid fruits that have been kept in vessels of galvanized iron. On this account, specimens intended for analysis in cases of supposed poisoning should never be placed in jars closed by zinc caps.

391. **Tests.** With **alkaline hydrates and carbonates,** solutions of zinc give a white precipitate, soluble in excess of the reagent; with **ammonium sulphydrate** or **sulphydric acid** in neutral or alkaline solutions, a white sulphide; with **potassium ferrocyanide,** a yellowish-white precipitate, insoluble in dilute hydrochloric acid.

CADMIUM.

Cd = 112.

392. **Occurrence, Preparation, Properties, etc.**—A comparatively rare metal, often accompanying zinc in its ores. Being more volatile than its associate, it accumulates during the first stages of the process of distilling zinc from its ores. It is a soft, white, tenacious metal, of specific gravity 8.6. It alters but little in the air at ordinary temperatures. Heated, it burns with formation of the oxide as a brown smoke. It dissolves with difficulty in sulphuric and hydrochloric acids, but readily in nitric acid.

393. **Cadmium Compounds.**—These are not very numerous or important. As the element is bivalent, they all have the general formula, CdR_2. The principal ones are: cadmium hydrate ($Cd(OH)_2$), cadmium oxide (CdO), cadmium chloride ($CdCl_2$), cadmium iodide (CdI_2), cadmium sulphate ($CdSO_4$) and cadmium sulphide (CdS). The latter is found native in the mineral Greenockite.

IRON GROUP.

Chromium = 52.4. **Manganese** = 54.8. **Iron** = 56.

394. The elements of this group form two series of compounds, due to a peculiar variation of the quantivalence. In one series, a single atom enters into combinations as a dyad, as in $Fe''Cl_2$; in the other, two atoms combine to form a hexad group, as in $(Fe_2)^{vi}Cl_6$. In this hexad condition they form acid radicals with oxygen, and form chromates, manganates and ferrates.

CHROMIUM.

$Cr = 52.4.$

395. **Occurrence, Preparation and Properties.**—This metal most commonly occurs in chromite, or chrome iron ore, a ferroso-chromic oxide; also, rarely, as lead chromate. It may be isolated with difficulty from its oxide by reducing with charcoal; or from the chloride by reducing with zinc. It is a hard, glistening, steel-gray metal, magnetic at low temperatures; sp. gr. 6.8; oxidizes only at a red heat, and is soluble in hydrochloric acid and strong alkalies.

396. **Chromic Anhydride—Chromic Trioxide** ($Cr^{vi}O_3$).—This is sometimes improperly called chromic acid. It is prepared by adding one and one-half parts strong sulphuric acid to one part of concentrated solution of potassium dichromate. When the solution cools, splendid saffron-colored needles of the trioxide crystallize out, which may be dried on a porous tile. It is a powerful oxidant, igniting alcohol if the latter be poured upon it. It is used in medicine as a caustic, forming a superficial eschar.

397. **Chromic Oxide—Chromium Sesquioxide—Green Oxide** ($Cr_2^{vi}O_3$), is obtained by calcining a mixture of starch and potassium dichromate. Thus prepared it is a green powder, insoluble in water, acids or alkalies, and fusible with difficulty. When fused with alkaline hydrates or nitrates, it forms chromates of these metals.

This oxide may play either a positive or negative rôle, depending upon the oxide with which it unites. For example, with the strongly negative sulphuric oxide, it forms chromium sulphate ($Cr_2(SO_4)_3$); while with calcium or magnesium oxide, calcium or magnesium chromites ($CaCr_2O_4$ or $MgCr_2O_4$) are obtained. The best known of these compounds is ($FeCr_2O_4$) ferrous chromite, or native chrome iron.

398. **Chromous Oxide** ($Cr''O$).—This exists only as a hydrate, produced by precipitating chromous chloride by potassium hydrate. It acts as a basic oxide, yielding chromous salts.

399. **Chromic Acid** (H_2CrO_4).—This cannot be isolated, but by solution of chromium trioxide in water, an acid liquid is obtained containing chromic acid. This acid decomposes on evaporation.

The best known of the salts of this acid are potassium chromate and dichromate (K_2CrO_4 and $K_2Cr_2O_7$). The last of these possesses the properties of the trioxide, but in a milder degree. It is sometimes used as an escharotic, but much more frequently as an oxidizing agent, mixed with sulphuric acid. None of the other preparations of chromium are used in medicine. Internally, in large doses, it acts as an irritant poison. Antidotes—mild alkalies, demulcents. A form of chronic poisoning sometimes affects workmen who handle it.

400. **Chlorides.**—Two chlorides are known: **Chromous Chloride** ($CrCl_2$), a white, crystalline solid, dissolving in water to form a blue solution; and **chromic chloride** (Cr_2Cl_6), occurring in large red crystals, insoluble in water, unless a trace of the chromous chloride be present, when it dissolves readily. If it be subjected to a prolonged boiling, it finally dissolves, forming a green solution containing a hydrate. An oxychloride is also known.

401. **Chromic Sulphate** ($Cr_2(SO_4)_3$) is obtained by dissolving chromic oxide in sulphuric acid; upon slowly evaporating, it crystallizes with twelve molecules of water.

Chromium salts form two series, the one green and the other violet. The alkaline hydrates throw down a bluish-green hydrate from the green salts, and a violet from the violet.

Chromium sulphate exists as a violet crystalline solid, and as a green amorphous solid. With the alkaline sulphates, chromium sulphate forms double salts—the chromium alums. (See Alums.)

402. **Potassium Chromate** (K_2CrO_4) is obtained by adding a solution of potassium hydrate to one of potassium dichromate. It forms large, yellow, rhombic crystals, isomorphous with potassium sulphate (K_2SO_4).

403. **Potassium Dichromate** ($K_2Cr_2O_7$), commercially known as the **red chromate of potash** and often called the acid potassium chromate, is obtained by igniting pulverized **chromite** (Cr_2O_3FeO) with potassium carbonate and nitrate, forming potassium chromate and ferric oxide. The potassium

chromate is dissolved out with water and nitric or acetic acid added to the solution, from which potassium bichromate crystallizes. It forms large, red prisms, soluble at ordinary temperatures in ten parts of water. When it is warmed with sulphuric acid, oxygen escapes, and chromic acid and potassium chrome alum are produced. This mixture is employed in laboratories for oxidizing purposes.

404. **Toxicology.**—The chromates, especially potassium dichromate, are irritant poisons. They are also liable to produce a form of chronic poisoning in workmen handling them, characterized by ulceration of the septum of the nose, and of the skin. The most prominent symptoms in acute poisoning are vomiting, epigastric pain, cramps, excessive thirst and collapse. The treatment consists in the use of emetics, followed by magnesium carbonate in milk.

MANGANESE.

$Mn = 55.$

405. **Occurrence.** Manganese is found widely distributed in nature. It occurs native in meteorites. Its most common ores are **pyrolusite** (MnO_2), **hausmanite** (Mn_3O_4), **braunite** (Mn_2O_3), **manganite** ($Mn_2O_3H_2O$), and **rhodochrosite** ($MnCO_3$).

406. **Preparation and Properties.** It is obtained in the metallic condition by heating its oxides with charcoal, similar to the smelting of iron. A grayish-white, brittle metal, very hard, and fusing with great difficulty; sp. gr. 7.2. Like the elements iron and chromium, it forms three series of compounds; the manganous (MnR_2), manganic (Mn_2R_4), and the derivatives of manganic acid, called **manganates.**

407. **Manganous Compounds.** In these the metal is diatomic. These derivatives are the most stable and constitute the most common of the manganese salts. They resemble the **ous** salts of iron and chromium, with which they are isomorphous. **Manganous oxide** (MnO) results from ignition of the carbonate with exclusion of air. It is a greenish, amorphous powder, readily oxidizing in the air to Mn_3O_4. **Manganous hydrate** ($Mn(OH_4)_2$) is formed by adding alkaline hydrates to manganous solutions, as a reddish-white precipitate, which, exposed to the air, oxidizes to manganic hydrate, and turns brown in color.

408. **Manganous Salts.—Manganous Chloride** ($MnCl_2$) occurs in rose-colored tabular crystals, which decompose on

drying, with separation of hydrochloric acid. **Manganous Sulphate** ($MnSO_4$) crystallizes at ordinary temperatures with $5H_2O$. With the alkaline sulphates it forms double salts; *e. g.*, $MnSO_4.K_2SO_4 + 5H_2O$. **Manganous Carbonate** ($MnCO_3$) is precipitated from manganous solutions by alkaline carbonates as a white powder, turning brown on exposure. **Manganous Sulphide** (MnS) occurs in nature as alabandite, or manganese blende, and is precipitated from manganous solutions by alkaline sulphides as a flesh-colored hydrate ($MnS + H_2O$). In the air it also becomes brown.

409. **Manganic Compounds.**—These are isomorphous with, and very closely resemble the ferric, chromic and aluminic compounds. They are not so stable, however, being easily reduced to the manganous state. In them manganese is a tetrad.

410. **Manganese Dioxide — Manganic Peroxide — Black Oxide of Manganese** (MnO_2), occurs native as the mineral **pyrolusite,** the principal ore of manganese, in steel-gray or brownish, imperfectly crystallized masses.

When gently heated, it yields oxygen; at a red heat, it yields more oxygen and forms manganous-manganic oxide.

$$3MnO_2 = Mn_3O_4 + 2O.$$

It gives off oxygen when heated with sulphuric acid, and forms manganous sulphate. With hydrochloric acid, it yields manganous chloride, water and chlorine.

$$MnO_2 + 4HCl = MnCl_2 + 2H_2O + Cl_2.$$

In cold hydrochloric acid, it dissolves without setting chlorine free, as $MnCl_4$ is probably formed, which, upon heating, breaks up into $MnCl_2$ and Cl_2. From this it would appear that, in the dioxide, manganese is a tetrad.

Manganic Oxide (Mn_2O_3) is a black powder, produced by igniting the manganese oxides in a current of oxygen.

Manganic Hydrate (Mn_2OH_6) is precipitated from manganic solutions by ammonium hydrate as a flesh-colored precipitate, rapidly turning brown.

Manganous-manganic Oxide ($Mn_3O_4 = MnO,Mn_2O_3$).—This is formed by the ignition of all the oxides in the air; it is isomorphous with **Magnetite** (Fe_3O_4).

Manganic Sulphate ($Mn_2(SO_4)_3$) is produced by the action of sulphuric acid upon manganic hydrate.

411. **Manganates and Permanganates.**—The derivatives of manganic acid ($H_2MnO_4 = MnO_2(OH)_2$) are analogous to

those of ferric (H_2FeO_4), chromic (H_2CrO_4), and sulphuric acid (H_2SO_4). In these derivatives, manganese is a hexad. The manganates are of little permanency and little used.

Potassium Manganate (K_2MnO_4) is a rare substance, isomorphous with potassium sulphate or chromate, and is very readily converted into potassium permanganate.

412. **Potassium Permanganate** ($KMnO_4$) is precipitated from solutions of potassium manganate by acids, in dark-red, rhombic prisms, isomorphous with potassium chlorate. This salt has active oxidizing properties, and is very largely used for oxidizing and destroying organic substances. It also converts ferrous into ferric salts, and is used for the quantitative estimation of ferrous salts.

IRON.

Fe = 56.

413. **Occurrence.**—This metal, which is of so great practical importance, is distributed very widely in nature. It occurs native upon the earth's surface only as meteorites.

The ores from which iron is obtained are numerous; the most important are: **magnetite** (Fe_3O_4), **hæmatite** (Fe_2O_3), **limonite** (ferric hydrate) and **siderite** ($FeCO_3$). These are almost the only ores used for the manufacture of iron; the sulphur ores are not adapted for this purpose.

414. **Preparation.**—In some cases the ore is first roasted, to get rid of water, carbon dioxide, sulphur, etc. The next step consists in the extraction of the iron from the ores, in which it exists as oxide; this is accomplished by reduction with carbon, at a glowing heat. This reduction is effected in a **blast furnace,** of which the interior has the shape of a double cone. It is about fifty or sixty feet high by fifteen feet wide at its widest part, is built of the most infusible fire brick, and inclosed in solid masonry. It is filled at the top with alternate layers of coal, broken ore (either native or previously roasted), and fluxes in the form of limestone or silicates. These fluxes facilitate the melting together of the reduced iron, and furnish a liquid slag. The air necessary for combustion, usually heated to a high temperature beforehand, is forced into the bottom of the furnace, through pipes, by blowers or fans. The metal is drawn off at the bottom. In the lower part of the furnace, carbon dioxide is produced from the oxygen of the air and the coal; and higher up, carbon monoxide is produced, which acts upon the oxide of iron, reducing it to the metallic state. As the reduced

iron sinks, it comes into contact with the coal, takes up a small quantity of carbon, and forms cast-iron, which, on further sinking, fuses, and is drawn off into moulds made in sand, to form **pig iron.** The earthy impurities of the ores remaining in the furnace unite with the fluxes, fuse in the intense heat, and are drawn off as **slag.** The pig iron is then subjected to the **puddling process,** by which it is more completely freed from carbon and slag, and **wrought iron** results. This process is usually carried on in reverberatory furnaces with a free supply of air, while the molten mass is being thoroughly stirred. The greater part of the carbon is in this way burned into carbon monoxide, and the silicon, sulphur and phosphorus oxidized. **Steel** was formerly prepared from wrought iron only, by cementation, or heating wrought iron in boxes packed in leather shavings or with charcoal. At present it is chiefly prepared directly from cast or pig iron by the method invented by Bessemer in 1850. This process consists in blowing air, under high pressure, into a mass of molten cast-iron, until the carbon has been consumed, when spiegeleisen, containing a known quantity of carbon, is added, to give the proper amount of carbon. **Pure iron** is obtained by heating ferric oxide in a current of hydrogen; this is the **ferrum redactum** (U. S., Br.).

415. **Properties.**—Pure iron is soft, fuses at about 1600°C. (2912°F.), and has a specific gravity of 7.25 to 7.9. Iron is not affected by dry air at ordinary temperatures; in moist air, it covers itself with a thin layer of ferric hydrate, known as **rust.** Heated strongly in the air, it becomes coated with a layer of ferrous-ferric oxide (Fe_3O_4), which is readily loosened, forming the blacksmith's scales; at a red heat, it decomposes water, with the formation of ferrous-ferric oxide, and the liberation of hydrogen.

$$3Fe + 4H_2O = Fe_3O_4 + 4H_2.$$

In oxygen it burns with an intense, scintillating light. If brought into contact with a magnet, iron becomes magnetic. Tempered steel is the only form, however, that retains the magnetism.

Iron unites directly with chlorine, bromine, iodine, sulphur, and the members of the phosphorus group, except nitrogen. It dissolves readily in hydrochloric and sulphuric acids, with evolution of hydrogen. In dilute nitric acid, it dissolves with separation of nitric oxide. Concentrated nitric acid, however, renders it **passive,** when it is no longer attacked by the dilute acid, until the passive condition is destroyed by contact with silver, platinum or copper, or by heating to 40° C. (104° F.).

416. **Ferrous Compounds.**—These are formed by dissolving iron in an acid, or by the reduction of ferric salts.

$$Fe_2Cl + Zn = (FeCl_2)_2 + ZnCl_2.$$

They are usually of a green color in the hydrous state; exposed to the air, they oxidize to ferric salts.

$$2FeO + O = Fe_2O_3.$$

417. **Ferrous Chloride** ($FeCl_2$) is formed when iron is dissolved in hydrochloric acid. It crystallizes in green, monoclinic prisms, containing four molecules of water. Exposed to the air, they deliquesce and oxidize, forming ferric chloride and an oxychloride. The anhydrous chloride is formed by passing hydrochloric acid gas over iron that is heated to redness, as a volatile, yellowish-white, very soluble solid.

418. **Ferrous Iodide—Ferri Iodidum** (Br.), (FeI_2), is obtained in solution by adding an excess of iron to iodine suspended in warm water until the solution is pale green.

419. **Ferrous Oxide** (FeO) is a black powder, produced by the reduction of ferric oxide by carbon. It easily oxidizes again.

Ferrous Hydrate ($Fe(OH)_2$) is precipitated from ferrous solutions by alkali hydrates as a white powder. It also oxidizes readily, becoming green, and then brown.

420. **Ferrous Sulphate — Protosulphate of Iron — Green Vitriol — Copperas — Ferri Sulphas** (U. S., Br.), ($FeSO_4 7H_2O$), is obtained pure by dissolving iron in dilute sulphuric acid. For commercial use it is obtained from **pyrites** (FeS_2) by oxidation, and as a by-product in other processes. It forms oblique rhombic prisms. At a red heat, it decomposes into ferric oxide, and sulphur di- and trioxides. On this property is based the production of fuming, or **Nordhausen sulphuric acid.**

Green vitriol has an extended use in the arts. Among other uses, it is employed in the manufacture of ink, and as a mordaunt in dyeing.

Ammonio-ferrous Sulphate ($Fe(NH_4)_2.(SO_4)_2.6H_2O$) is also known. It is more stable than ferrous sulphate.

421. **Ferrous Carbonate** ($FeCO_3$) exists in the mineral, **siderite.** It may be obtained by adding sodium carbonate to ferrous solutions. It is readily oxidized to ferrous hydrate on exposure to the air; is insoluble in pure water, but soluble in water containing carbon dioxide, and is, therefore, present in many natural waters.

The **ferri carbonas saccharatum** is ferrous carbonate to which sugar has been added to prevent decomposition, and is prepared by mixing solutions of ferrous sulphate and sodium bicarbonate and adding sugar; the mixture is then evaporated to dryness. It is a greenish-gray powder.

422. **Ferrous Phosphate—Triferrous Phosphate** ($Fe_3(PO_4)_2$), is a white precipitate formed by adding sodium phosphate to a solution of a ferrous salt. It turns blue on exposure to the air, a part being converted into ferric phosphate. The **ferri phosphas** (U. S., Br.) is a mixture of these two salts. It is insoluble in water; slightly soluble in water containing carbonic dioxide or acetic acid. A soluble or acid phosphate exists in the shops. A phosphate of iron, that turns blue on exposure to the air, exists in the lungs, phthisis, in bones which have been buried for some time, and occasionally in pus.

423. **Ferrous Sulphide—Protosulphide** (FeS), may be obtained, first, by fusing a mixture of sulphur and iron filings at ordinary temperatures, although the union will often occur slowly; second, by precipitation of a ferrous salt with alkaline sulphides. The first method forms brownish, brittle, fusible masses; the latter yields a black powder. Ferrous sulphide is not decomposed by heat, but is decomposed by sulphuric acid, with formation of ferrous sulphate and hydrogen sulphide. It occurs in the fæces of persons taking chalybeate waters and preparations of iron.

424. **Ferrous Lactate—Ferri Lactas** (U.S.), ($Fe(C_3H_5O_3)$), is obtained by dissolving iron filings in lactic acid. It forms light yellow crystals, soluble in water, insoluble in cold alcohol.

425. **Ferrous Oxalate—Ferri Oxalas** (U. S.), ($Fe(C_2O_4)$), is made by dissolving iron in a solution of oxalic acid. A bright yellow, crystalline powder, slightly soluble in hot water.

426. **Ferrous Tartrate** ($FeC_4H_4O_6$) is formed by dissolving iron in a hot, strong solution of tartaric acid.

427. **Ferric Compounds—Ferric Oxide—Sesquioxide of Iron** (Fe_2O_3), exists in nature as hæmatite, and may be formed by heating the oxygen compounds of iron in the air. On a large scale, it is obtained by distilling ferrous sulphate, which first turns white, from loss of water; then yellow, owing to the formation of an oxyhydrate, and finally to a brick-red, ferric oxide. It is used as a polishing material, under the names of **colcothar, red crocus, jewelers' rouge, caput mortuum,** or **Venetian red.**

Ferrous-Ferric Oxide ($Fe_3O_4 = FeOFe_2O_3$) occurs native as **magnetite.** It may be obtained by conducting steam over ignited iron. It constitutes the natural magnets.

428. **Ferric Hydrate — Ferri Oxidum Hydratum** (U. S., Br.), ($Fe_2(OH)_6$), is a voluminous, reddish-brown, gelatinous mass, precipitated by alkaline hydrates from ferric solutions. When dried at 100°C. (212°F.), it loses $2H_2O$. Freshly precipitated ferric hydrate is soluble in a solution of ferric chloride or acetate, and if such a solution is dialysed, the iron salt diffuses, leaving the pure ferric hydrate on the dialyser. The **dialysed** iron so obtained is coagulated by heat, acids or alkalies, into a jelly-like mass. It is a good antidote in arsenic poisoning.

429. **Ferric Chloride—Sesquichloride of Iron—Perchloride of Iron—Ferri Chloridum** (U.S.), (F_2Cl_6), may be obtained anhydrous in volatile, deliquescent plates by heating iron in chlorine gas. It may be formed in solution by dissolving iron in hydrochloric acid and adding a little chlorine water or nitric acid; or by dissolving the oxide or hydrate in hydrochloric acid; or by the action of chlorine on a solution of ferrous chloride. To obtain the solid, it is only necessary to evaporate and crystallize.

It is prepared in pharmacy by the last method. It forms yellow, crystalline masses or rhombic plates, readily soluble in water, alcohol or ether.

The **Liq. ferri chloridi** (U. S.), or **Liq. ferri perchloridi** (Br.), is an aqueous solution containing an excess of acid.

The U. S. P. preparation contains 37.8 per cent. of ferric chloride.

The **Tinct. ferri chloridi** (U. S.) is the same, diluted with alcohol, and contains also ethyl chloride and ferrous chloride.

430. **Ferric Sulphate** ($Fe_2(SO_4)_3$) is obtained by dissolving the oxide in sulphuric acid. It remains, after evaporating the solution, as a white mass, which dissolves readily in water, forming the **Liquor ferri tersulphatis** (U. S.). This solution is a dark, reddish-brown liquid, having an acid, styptic taste. Another sulphate, which is basic, is formed by treating ferrous sulphate, 77 parts, with nitric acid, and evaporating, after adding 7 parts sulphuric acid.

This, in solution, is the **Liq. ferri subsulphatis**, or **Monsel's solution.**

431. **Ferric Alum—Ammonio-ferric Alum—Ferri et Ammoniæ Sulphas** (U.S.P.), $(NH_4)_2Fe_2(SO_4)_4.24H_2O$, is prepared by adding a solution of ammonium sulphate to a solution of

ferric sulphate, and evaporating down and allowing it to crystallize. The crystals are colorless or pale amethyst, regular octahedra, soluble in 3 parts of water at 15°C. (59°F.). The solution has an acid reaction and an astringent taste, but not so astringent as ferric sulphate. It gives the reactions for ferric iron, ammonia and sulphates. It is employed in medicine as an astringent, both internally and locally.

432. **Ferric Nitrate** ($Fe_2(NO_3)_6$) is formed, together with ferrous nitrate, by dissolving iron in nitric acid. The **Liq. ferri nitratis** (U. S.), or **Liq. ferri pernitratis** (Br.), is an aqueous solution of ferric nitrate. It crystallizes in rhombic prisms with $18H_2O$, or in cubes with $12H_2O$.

433. **Scale Compounds of Iron.**—These are certain salts of iron, mostly with organic acids, which do not crystallize readily, but are put into the market in the form of thin scales. They are prepared by evaporating their solution to a thick, syrupy consistence, spreading upon glass plates, drying, and then detaching the thin scales from the glass. They are all used in medicine.

Ferri Citras (U. S.)—citrate of iron—is prepared by the action of citric acid upon soluble salts of iron, usually the sulphate. After its aqueous solution has been evaporated upon glass, it forms beautiful, thin, transparent scales, of a garnet-red color, slowly soluble in cold, but freely in hot water, and possessing a mild, chalybeate taste.

Ferri et Ammoniæ Citras (U. S., Br.)—citrate of iron and ammonia—is formed by treating a solution of citrate of iron with ammonium hydrate, and evaporating at a temperature that should not exceed 38°C. (100.4°F.).

It also forms garnet-red scales, which are readily and wholly soluble in water, forming a solution that is neutral to litmus paper and slightly styptic in taste.

Ferri et Ammoniæ Tartras (U. S.)—tartrate of iron and ammonia—a double salt, is formed by the action of tartaric acid upon ferric and ammonic hydrates. Upon evaporation of its solution, garnet-red scales remain, which are slowly soluble in water. Their solution is neutral to test paper, and is of a sweetish, rather pleasant taste.

Ferri et Potassæ Tartras (U. S.)—**Ferrum Tartaratum** (Br.)—potassio-tartrate of iron—may be obtained by dissolving ferric hydrate in a solution of acid tartrate of potassium, and evaporating on glass. It forms ruby-red plates, having about the same properties as the ammonio-tartrate.

Ferri et Quiniæ Citras (U. S., Br.)—citrate of iron and quinine—contains citric acid, ferric hydrate and quinia citrate. It forms transparent scales of a greenish tint, slowly soluble in cold water, but freely in hot water, forming bitter, slightly styptic solutions.

Ferri et Strychniæ Citras closely resembles the citrate of iron and ammonia in appearance, but has a bitter taste, and gives a white precipitate with ammonium hydrate. It is prepared by adding a solution of strychnine citrate to a solution of ferric citrate, evaporating on a water bath to a syrup, and drying on glass.

Ferri Phosphas (U. S., Br.), ($Fe_2(PO_4)_2$).—Phosphate of iron occurs as the result of a double decomposition between ferric sulphate and sodium phosphate. It forms a bright, slate-colored powder, insoluble in water, but soluble in acids.

Ferri Pyrophosphas (U. S.), ($Fe_4(P_2O_7)_3$).—Pyrophosphate of iron is formed by adding a solution of sodium pyrophosphate to a solution of a ferric salt. It does not crystallize, but forms scales upon evaporating its solution. These are thin, apple-green in color, turn dark on exposure to the air, and are soluble in water, but not in alcohol. Ammonium hydrate produces no precipitate in its solutions, but sodium hydrate does. The officinal salt contains 48 per cent. of anhydrous ferric pyrophosphate.

NICKEL GROUP.

Nickel = 59. **Cobalt** = 59.

434. These elements are often included with the members of the iron group, to which they bear a certain resemblance, but from which they differ in not forming, so far as is known, compounds similar to the ferrates, chromates or manganates. Furthermore, no compounds of nickel or cobalt analogous to the salts of chromic or manganic acids are known.

NICKEL.

435. **Occurrence, Preparation and Properties.**—This metal is found native in meteorites. Its most common ores are **Niccolite** (Ni_2As_2) and **Gersdorffite** ($(NiS_2)_2As_2$). These ores of nickel, however, usually also contain cobalt, and the cobaltous ores are commonly nickel bearing also. The separation of nickel from its ores is a very complicated process, and for an account of it, the reader is referred to works on metallurgy. It may be prepared chemically pure by igniting its oxalate or car-

bonate in a stream of hydrogen. This metal is silver white, tenacious and very lustrous; sp. gr. 9.1. It is attracted by the magnet. It does not alter in the air, but dissolves in the mineral acids, especially nitric.

436. **Compounds.**—The following are the most common of the nickel compounds; they are all **ous** compounds, having the general form NiR_2, and all possess a green color: **nickelous hydrate** ($Ni(OH)_2$), **nickelous chloride** ($NiCl_2 + 6H_2O$), **nickelous cyanide** ($Ni(CN)_2$), **nickelous sulphate** ($NiSO_4 + 7H_2O$) and **nickelous sulphide** (NiS). **Nickelic oxide** (Ni_2O_3) and **hydrate** ($Ni_2(OH)_6$) exist, and are similar to the corresponding cobalt compounds. Nickel is used largely in certain alloys and for electro-plating.

COBALT.

Co = 59.

437. **Occurrence, Preparation, Properties, etc.**—**Smaltite** ($CoAs_2$) and **cobaltite** ($CoAs_2 . CoS_2$) are the most commonly occurring native ores of cobalt. It is prepared in the same manner as nickel. It is a reddish-white metal, tenacious, and fusible with great difficulty; sp. gr. 8.9. Its other properties are very similar to those of nickel.

438. **Cobalt Compounds** are also chiefly **ous**, corresponding to the general form CoR_2. Those containing water have a reddish color; the anhydrous compounds are blue.

The cobaltous compounds are, **cobaltous chloride** ($CoCl_2$), **cobaltous hydrate** ($Co(HO)_2$), **cobaltous sulphate** ($CoSO_4 + 7H_2O$), **cobaltous nitrate** ($Co(NO_3)_2 + 6H_2O$), and **cobaltous sulphide** (CoS).

The cobaltic compounds are, **cobaltic oxide** (Co_2O_3), and **cobaltous-cobaltic oxide** ($Co_3O_4 = Co_2O_3CoO$). The latter corresponds to magnetite (Fe_3O_4).

LEAD GROUP.

Lead, Pb = 207.

439. **Occurrence.**—The most abundant ore found native is **Galena**, or **Galenite** (PbS). Other ores are **Cerussite** ($PbCO_3$), **Crocoisite** ($PbCrO_4$), **Wulfenite** ($PbMoO_4$) and **pyromasphite** ($Pb_3(PO_4)_2$).

440. **Preparation.**—For this purpose **galenite** is almost exclusively employed. The ore is first roasted in the air, by

which a portion of the lead sulphide is converted into oxide and another part into sulphate.

$$\begin{aligned} &PbS + 3O = PbO + SO_2 \\ \text{and}\quad &PbS + 4O = PbSO_4. \end{aligned}$$

These two products are then strongly heated in a reverberatory furnace, when they react as follows:—

$$\begin{aligned} &2PbO + PbS = 3Pb + SO_2 \\ \text{and}\quad &PbSO_4 + PbS = 2Pb + 2SO_2 \end{aligned}$$

If the galena contain much silver, this is separated by crystallization and cupellation.

441. **Properties.**—Lead is a bluish-white metal, brilliant upon freshly cut surfaces, but soon tarnishes. It is soft and pliable, but not very malleable or ductile; specific gravity 11.37. It fuses at 334° C. (633° F.). If melted, it does not, on cooling, return to its original volume. It is a poor conductor of electricity, but a better conductor of heat.

When exposed to the air, it oxidizes slightly. It is not acted upon by pure water deprived of air, but, by the contact of air and water, it oxidizes to the hydrate ($Pb(OH)_2$), which dissolves slightly in water. If the water contains carbon dioxide, carbonates or sulphates, very little lead goes into solution, but it is coated with an insoluble layer of lead carbonate or sulphate. (If carbon dioxide be in excess, as in soda water, the carbonate formed is somewhat soluble in the water).

The solvent action of water upon lead is increased, however, by the presence of nitrates and nitrites. These facts are of great practical importance, as lead pipes are very frequently employed for conducting water.

Sulphuric and hydrochloric acids have but little effect on lead, especially if cold, owing to the insolubility of its sulphate and chloride. Nitric acid dissolves it readily. Zinc, tin and iron precipitate this metal from its solution.

There are several useful alloys of lead. Alloyed with an equal part of tin, it fuses at 186° C. (366.8° F.) and is used for soldering (soft solder). Type-metal is an alloy of four or five parts of lead and one of antimony; the proportions vary considerably.

442. **Lead Oxide—Protoxide—Massicot—Litharge—Plumbi Oxidum** (U. S., Br.), (PbO), is prepared by heating lead, its carbonate or nitrate, in the air. If it fuses, it constitutes **litharge**; if not, **massicot.** The former is a reddish-yellow or brown mass of rhombic scales; the latter is a yellow,

amorphous powder, differing from litharge in color and texture, but not in composition.

Lead oxide has strong basic properties. It absorbs carbon dioxide from the air and imparts an alkaline reaction to water, in which it dissolves as hydrate. Like other strong bases, it saponifies fats and forms lead soaps, as **lead plaster.** It dissolves readily in nitric or hot acetic acid, with formation of nitrate or acetate of lead. It fuses at a red heat. If fused in earthen crucibles, it forms a silicate, and thus perforates the crucible. When heated to 300° C. (572° F.) in contact with air, it is oxidized to **minium**, or a bright red powder (red lead).

443. **Plumboso-plumbic Oxide—Minium—Red Lead** (Pb_3O_4), or ($2PbO.PbO_2$), is prepared as already stated, by roasting litharge at a temperature of 300° C. (572° F.), and is used as a pigment and in the manufacture of glass. Its composition is probably expressed by the formula Pb_3O_4; or, as one molecule of the dioxide combined with two of the monoxide; or, as the lead salt of plumbic acid.

It is a brilliant red powder, of specific gravity 8.62. When strongly heated, or subjected to the action of reducing agents, it is converted into litharge. Nitric acid dissolves the monoxide, leaving the dioxide, and changing its color to brown.

As occurring in commerce, it is frequently contaminated with oxides of iron or brickdust. It should dissolve in dilute nitric acid to which a little sugar has been added.

444. **Lead Dioxide—Peroxide of Lead—Puce Oxide of Lead—Binoxide of Lead—Plumbic Anhydride**—may be prepared by dissolving the monoxide out of minium with dilute nitric acid, or by the action of chlorine upon lead carbonate suspended in water.

It is a dark, reddish-brown powder, insoluble in water; specific gravity 8.903 to 9.190. Heat drives off half its oxygen, converting it into monoxide. It is a valuable oxidizing agent.

445. **Plumbic Acid** is formed as crystalline plates at the positive electrode when alkaline solutions of the lead salts are subjected to electrolysis.

With the alkaline hydrates, lead dioxide dissolves to form well defined but unstable **plumbates. Potassium** plumbate may be obtained in cubic crystals by dissolving the hydrate in potassium hydrate, and cooling the solution. It is decomposed by water.

446. **Lead Chloride** ($PbCl_2$) separates as a white precipitate when hydrochloric acid is added to a solution of a lead salt. It

is nearly insoluble in cold water, but dissolves in thirty parts of hot water, from which solution it crystallizes on cooling, in white, shining needles. At a red heat, it fuses to a horn-like mass.

447. **Lead Iodide—Plumbi Iodidum** (U. S., Br.), (PbI_2), is precipitated from lead solutions by potassium iodide, as a bright yellow, crystalline powder. It is practically insoluble in cold water, but more soluble in boiling water, from which it crystallizes on cooling, in beautiful, gold-colored, glistening crystals. Exposed to light and moisture, it decomposes, with liberation of iodine.

448. **Lead Nitrate—Neutral Lead Nitrate—Plumbi Nitras** (U. S., Br.), ($Pb(NO_3)_2$), is obtained by dissolving lead or its oxides in excess of nitric acid. It forms anhydrous, octahedral crystals, soluble in 1.98 parts of water at 17.5°C. (63.5°F.), and in 0.7 parts at 100°C. (212°F.). At a red heat, it melts and is decomposed into PbO,NO_2 and oxygen, and at a higher temperature, into PbO.

449. **Lead Sulphate** ($PbSO_4$) occurs in the mineral **Anglesite** in rhombic crystals, isomorphous with barium sulphate. It is produced by the double decomposition between a sulphate and a soluble lead salt. It is insoluble in water, but readily soluble in concentrated sulphuric acid. The commercial acid always contains it.

450. **Lead Carbonate—Plumbi Carbonas** (U. S., Br.), ($PbCO_3$), occurs as **Cerussite.** It may be produced by double decomposition between a lead salt and a carbonate, or by passing carbon dioxide through a solution of a lead salt. **White lead** is usually prepared commercially by treating thin sheets of lead with acetic acid and then exposing the acetate to carbon dioxide. The lead, rolled into sheets, is placed in earthen jars containing a small quantity of vinegar at the bottom, but not in contact with the lead. Great numbers of the jars after being thus charged are buried in stable manure or spent tan bark. By the decomposition of the bark or manure, considerable carbon dioxide and heat are produced. The heat volatilizes a portion of the vinegar, which, acting upon the lead, produces the acetate ($Pb(C_2H_3O_2)_2$). The carbon dioxide acting upon the acetate converts it into acetic acid, which acts upon a fresh portion of lead, and a basic or hydro-carbonate of lead.

$$2PbCO_3 . PbO_2H_2 \text{ or } PbO—HCO_3—Pb—CO_3—PbO—H.$$

After the lapse of a considerable time, the pile is taken down, the sheets are taken out, and the carbonate detached from them by passing them through rollers or by pounding. The white powder is then ground with oil and sent into the market as "**White Lead.**"

White lead is largely used in oil painting, forming a part of all but the darkest colors. As it is poisonous and is darkened by the action of hydrogen sulphide in the atmosphere, it is at present being more and more replaced by **zinc white** (ZnO) and **permanent white** ($BaSO_4$).

451. **Lead Sulphide** (PbS) occurs in the mineral **galena**. It is precipitated from solutions of lead salts by hydrogen sulphide or alkaline sulphydrates as a black powder. The native sulphide is bluish-gray and has a metallic lustre; sp. gr. 7.58. The sulphide obtained by precipitation has a sp. gr. of 6.924. It is insoluble in dilute acids.

452. **Lead Acetate—Salt of Saturn—Sugar of Lead—Plumbi Acetas** (U. S., Br.), ($Pb(C_2H_3O_2)_2$), is prepared by dissolving litharge in acetic acid; or by exposing lead in contact with acetic acid to the atmosphere, evaporating and crystallizing. It forms large, oblique rhombic prisms, having a sweetish, metallic taste. It dissolves in 1.5 parts of water and in 8 parts of alcohol, forming solutions which react acid upon test paper. On exposure to the air, the crystals effloresce upon the surface and are partly converted into carbonate. Several subacetates, as basic acetates, are known. The only one requiring mention is that having the formula $PbOH\ (C_2H_3O_2)2PbO$. This is the chief constituent of **Liq. plumbi subacetatis** (U. S., Br.), **or Goulard's extract**, which is obtained by boiling a solution of the neutral acetate with lead monoxide in fine powder. When exposed to the air, this solution becomes milky from the formation of lead carbonate.

453. **Lead Chromate** ($PbCrO_4$) is formed by precipitating lead nitrate with potassium chromate. Under the name of **chrome yellow**, it is used as a pigment. Very recently, its fraudulent use as an artificial coloring agent in manufactured food products has been discovered. It is insoluble in water, but soluble in alkalies.

454. **Physiological Action of Lead.** All of the compounds of lead that are soluble, and those that are themselves insoluble but that are readily convertible into soluble compounds by the action of air, water or the digestive fluids, are poisonous.

The chronic form of lead poisoning, **painter's colic**, is very common, and is produced by the continuous absorption of small quantities of the metal or its compounds, either by the skin, lungs, or stomach. Although metallic lead is inert, its absorption will cause symptoms of poisoning from its being converted within the body into poisonous compounds. Some of the methods by which it may be introduced are, the drinking of water that has been in contact with the metal; the use of food, tobacco, etc., that has been wrapped in tin-foil containing lead; the drinking of beer or other beverages that have been kept in pewter vessels; the handling of the metal or its alloys by artisans. Almost all of the commoner compounds of lead may give rise to the chronic poisoning. Probably the carbonate is the cause of more cases than any other lead compound in painters, artists, manufacturers of paint, and persons sleeping in freshly painted apartments.

Acute lead poisoning is comparatively rare and is not often fatal. It is generally caused by the ingestion of a single large dose of the acetate, subacetate, carbonate or red lead.

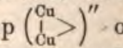
When it occurs, magnesium sulphate should be given, as it forms an insoluble lead sulphate.

If the metal be once absorbed, it is eliminated slowly, as it tends to become fixed by combination with the albuminoids of the body. This combination is rendered soluble by potassium iodide. It is eliminated by the urine, perspiration and bile.

Great caution is necessary in drawing conclusions, in case traces of lead are found in the body after death, on account of the many ways that it may be introduced.

COPPER GROUP.

Copper = 63.5. Mercury = 200.

455. Each of these elements forms two series of compounds. One series contains the group  $\left(\begin{matrix}\mathrm{Cu}\\ | \\ \mathrm{Cu}\end{matrix}\!>\right)''$ or $(Hg_2)''$ and the compounds are known as the cuprous and mercurous compounds, respectively; the other contains a single dyad atom, and is designated by the termination *ic*.

COPPER.

Cu = 63.5.

456. **Occurrence.**—This metal occurs in the free state in large masses, or crystallized in cubes and octahedra. It is found in the vicinity of Lake Superior, in China, Japan, Sweden, and in the Urals. Its most important ores are: cuprite (Cu_2O),

malachite and azurite (basic carbonates), chalcocite (Cu_2S), and chalcopyrite, or copper pyrites ($CuFe_2S$).

457. **Preparation.**—The mixed copper ores are first roasted in the air, by which process a portion of the copper sulphide is converted into oxide; this is then roasted with silica fluxes and carbon. By this process the iron sulphide is converted into silicate, and is drawn off with the slag. After several repetitions of this process, the so-called **copperstone** is obtained; this contains both the sulphide and oxide. By repeated roasting and heating, the copper oxide reacts upon the sulphide, and metallic copper results. Some poor ores are first treated with sulphuric acid, and the resulting sulphate is then treated with scrap iron, which precipitates the copper in the metallic state. Chemically pure copper is obtained by heating the pure oxide in a stream of hydrogen, or by electrolysis.

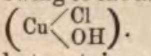
458. **Properties.**—Copper is a red metal by reflected light, while thin leaflets transmit a green light.

It is soft, ductile and tenacious; a good conductor of heat and electricity; specific gravity 8.914 to 8.952. In dry air it undergoes no change; but in moist air it gradually becomes coated with a thin layer of green basic carbonate. When heated, it oxidizes to black cupric oxide (CuO). Hot sulphuric, nitric and hydrochloric acids dissolve it, with liberation of sulphur dioxide, nitrogen dioxide, and hydrogen, respectively. With organic acids it forms soluble salts in the presence of air and moisture; hence, acid fruits should not be kept in copper vessels.

459. **Cuprous Compounds.**—These are very unstable, absorb oxygen, and are converted into cupric compounds.

If the formulæ $CuCl$, CuI, Cu_2O and Cu_2S are correct, copper in the cuprous compounds would appear, like silver, to be univalent. It has never been determined, however, whether these formulæ really express the true molecules. Copper compounds are not volatile, and we have no means of ascertaining the size of the molecule. As has already been stated, most chemists believe that in the cuprous compounds copper is bivalent, and that they contain the group Cu_2'', whose valence is always two.

460. **Cuprous Oxide — Suboxide, or Black Oxide** (Cu_2O), occurs in nature as **cuprite.** It is obtained artificially by boiling an alkaline solution of grape sugar and copper sulphate. It precipitates as a bright red powder. (Fehling's and Trommer's tests.) The hydrate ($Cu_2(OH)_2$) is precipitated by the alkalies, from hydrochloric acid solutions of Cu_2Cl_2, as a yellow powder.

461. **Cuprous Chloride—Subchloride, or Protochloride** ($CuCl$ or Cu_2Cl_2), is produced, together with cupric chloride, by igniting metallic copper in chlorine gas; by dissolving cuprous oxide in hydrochloric acid without contact of air; or by the action of many reducing agents upon solutions of cupric chloride. It is a heavy white powder, rapidly becoming green in the air, owing to the absorption of oxygen, and formation of cupric chloride  $\left(Cu\begin{matrix}\diagup Cl\\ \diagdown OH\end{matrix}\right)$. It dissolves in concentrated hydrochloric acid, but not in water. With carbon monoxide it forms a crystallizable compound. Its hydrochloric acid solution is used in gas analysis to absorb this gas.

462. **Cuprous Iodide** (Cu_2I_2) is precipitated, together with iodine, from soluble cupric salts, by potassium iodide.

$$2CuSO_4 + 4KI = 2K_2SO_4 + Cu_2I_2 + I_2.$$

On dissolving out the iodine with ether, the iodide is left as a gray insoluble powder.

463. **Cuprous Sulphide—Subsulphide, or Protosulphide** (Cu_2S), occurs in nature as **chalcocite,** as soft, fusible, gray crystals; also in many double sulphides, among which the most important is the double sulphide of copper and iron, or **copper pyrites.**

464. **Cupric Compounds—Cupric Oxide—Binoxide—Black Oxide** (CuO), is prepared by heating copper turnings to redness in the air, or by calcining the nitrate. It forms a black amorphous powder, readily reduced by heated charcoal, hydrogen, or the alkaline metals, to the metallic state. If heated in the presence of organic substances, it oxidizes them completely, and is thereby reduced to metal. It is used in organic analysis for this purpose.

465. **Cupric Hydrate** ($Cu(OH)_2$) is formed as a voluminous bluish-white precipitate, when sodium or potassium hydrate is added to a solution of a copper salt. When heated, even under water, it becomes dehydrated and changed to black cupric oxide.

Copper oxide and hydrate dissolve in ammonium hydrate, forming a dark-blue solution. This solution is often used as a solvent for cellulose, from which solution acids precipitate it again. (See Celluloid.)

466. **Cupric Chloride** ($CuCl_2$) is formed by dissolving cupric oxide or carbonate in hydrochloric acid. From aqueous solutions it crystallizes in bright green, rhombic needles with

$12H_2O$. It is readily soluble in water and alcohol. When heated, it parts with its water, and forms anhydrous cupric chloride, which at a red heat gives off chlorine.

$$2CuCl_2 + \text{heat} = Cu_2Cl_2 + Cl_2.$$

Cupric bromide resembles the chloride. The iodide is not known.

467. **Cupric Sulphate — Blue Vitriol — Blue Stone — Cupric Sulphas** (U. S., Br.), ($CuSO_4 5H_2O$), may be prepared, first, by roasting chalcocite; second, from the water of copper mines; third, by exposing copper moistened with dilute sulphuric acid, to the air; fourth, by dissolving copper in hot, concentrated sulphuric acid. It forms large, blue triclinic crystals, which dissolve in 2.71 parts of water at 19°C. (66.2°F.), and in 0.55 parts of water at 100°C. (212°F.). It loses four molecules of water at 100°C. (212°F), while the fifth separates above 200°C. (392°F.), leaving a white amorphous powder, which readily takes up water, and in so doing, resumes its blue color. Solutions of copper salts have a blue color, acid reaction and metallic, styptic taste.

Ammonium hydrate added to a solution of copper sulphate precipitates a bluish-white cupric hydrate, which dissolves in an excess of the alkali, forming a deep blue solution. Alcohol floated on this solution causes to separate long, right rhombic prisms having the composition $CuSO_4 4NH_3 H_2O$, which are very soluble in water. Their solution constitutes **cuprammonium sulphate,** and they exist mixed with other substances in the **cuprum ammoniatum** (U. S.).

468. **Copper Carbonates.**—The neutral carbonate ($CuCO_3$) is not known. When alkaline carbonates are added to solutions of copper salts, the basic carbonate separates as a green precipitate, having the formula $CuCO_3.2Cu(OH_2)$. This occurs in nature, especially in Siberia, as **malachite.** Another basic salt, **tricupric carbonate,** or **sesquicarbonate** of copper, is the beautiful blue **azurite.**

469. **Copper Arsenite—Scheele's green,** is prepared by adding a solution of sodium or potassium arsenite to a solution of a copper salt. It is a green powder, composed of copper arsenite and hydrate. It is insoluble in water, but soluble in ammonium hydrate or the mineral acids. It is exceedingly poisonous, but is often used as a pigment to color wall papers, toys, and even confectionery.

Schweinfurt Green — Mitis Green, or Paris Green ($Cu(C_2H_3O_2) + 3(As_2CuO_4)$), is the commonest and most dangerous of the cupro-arsenical pigments. It is prepared by adding a concentrated solution of cupric acetate to a boiling solution of arsenious acid. It is an insoluble, green, crystalline powder, decomposed by prolonged boiling in water, by aqueous solutions of the alkalies, and by the mineral acids. It is also soluble in ammonium hydrate.

470. **Copper Acetates.** The **diacetate** ($Cu(C_2H_3O_2)_2$) is formed by the decomposition of a solution of copper sulphate by lead acetate. It separates in large, bluish-green, prismatic crystals, with one molecule of H_2O, which it loses at 140°C. (284°F.) The dry salt when heated to 250°C. (482°F.), decomposes with liberation of glacial acetic acid.

Basic Acetates—Verdigris—Cupric Subacetas (U. S.), is a complex mixture of copper acetate and hydrate. It is prepared by exposing to the air piles composed of alternate layers of grape skins and copper plates, and after some time removing the bluish-green coating from the copper plates.*

471. **Copper Pigments.**—The most important are: **Brighton Green,** a mixture of copper acetate and chalk. **Brunswick Green,** orginally an impure chloride, but now generally a mixture of carbonate and chalk. **Mountain Green,** or **Mineral Green,** is a native green carbonate of copper, sometimes containing orpiment. **Neuwieder Green,** another name for mineral green or Schweinfurt green, mixed with gypsum or barium sulphate. **Green Verditer** is a mixture of the basic carbonate, oxide, and chalk.

472. **Physiological Action of Copper.**—Until recently, toxicologists were universally of the opinion that all the copper salts are poisonous. Of late, however, this has been considerably modified. Most of the copper compounds have an irritant, local action if brought into contact with the gastric or intestinal mucous membrane, causing vomiting of greenish matter, cramps, etc. On the other hand, there are numerous instances in which severe illness, characterized by nervous and other constitutional symptoms, has followed the use of food that has been in contact with imperfectly tinned copper vessels. Some such cases have proven fatal. It has been conclusively shown, however, that **pure** and non-irritating copper compounds may be taken in considerable quantity without any bad results, excepting vomiting.

Copper sulphate is frequently used as an astringent in medicine, and has been recommended in cholera and dysenteric troubles. This salt may be taken in considerable doses, with only an emetic effect. Cases of acute poisoning are not common, but many are recorded. Chronic poisoning is occasionally seen in those who work in copper, characterized by colicky pains, emaciation, impaired digestion, diarrhœa, and often a catarrhal cough. In most cases, there is a green line on the margin of the gums. Copper is very likely to become contaminated with arsenic; and it is possible that some of the cases of reported copper poisoning ought to be attributed to arsenic. The organic salts of copper seem to be more poisonous than the inorganic. Canned peas, pickles, and other fruits are often contaminated with copper, and the manufacturers have frequently been punished by fines; but there exists a difference of opinion as to the dangers of copper in such goods. As long, however, as there is a chance for doubt, sanitary authorities should prohibit its use.* The chemist must remember that most articles of food contain traces of copper.

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* The term verdigris is now often popularly applied to the green carbonates, hydrates, or salts of organic acids, which accumulate on the surface of copper.

The treatment of cases of irritant copper poisoning should consist in the exhibition of milk, white of egg, and other albuminoid substances, with which the copper salt may form an inert compound. Emesis should be induced if it has not taken place spontaneously.

MERCURY.

Hg = 200.

473. **Occurrence.**—Mercury occurs in nature principally as **cinnabar**, or, rarely, in the form of small particles scattered through rocks. It is found in Spain, Peru, China, Japan, California and Mexico.

474. **Preparation.**—The native sulphide, or cinnabar, is roasted in reverberatory furnaces, thus burning out the sulphur, and distilling off the mercury and condensing it. Or, it is sometimes simply heated with iron, which removes the sulphur and sets free the mercury, which distills over. Commercial mercury usually contains small quantities of other metals, owing to its great tendency to amalgamate with them. To remove these it is re-distilled, or treated with very dilute acids by pouring it in a thin stream into them. When pure, mercury pours from a glass surface without leaving a streak, and the single droplets retain their globular form, and do not form a tail or adhere to the glass.

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* This has been done in Brooklyn, N. Y., where one case of fatal poisoning has occurred from a pickle containing copper.

475. **Properties.**—Mercury is the only liquid metal at ordinary temperatures. Its specific gravity is 13.596. At 40° C. it solidifies, and crystallizes in octahedra. It is somewhat volatile at ordinary temperatures, and boils at 360° C. (680° F.). Its vapor has a density of 100; specific gravity 6.97 (Air = 1). Its molecular weight, therefore, is 200; and as its atomic weight is also 200, its molecule, like that of cadmium, is composed of one atom. If pure and at ordinary temperatures, it is not altered in the air; at a temperature near the boiling point, it is coated with a thin film of mercuric oxide. Hot sulphuric acid converts it into mercuric sulphate, with evolution of sulphur dioxide. It dissolves readily in dilute nitric, but not in hydrochloric acid.

Mercury dissolves all metals but iron, to form **amalgams.** Tin amalgam is used for coating mirrors.

476. **Mercurous Compounds.—Mercurous Chloride—Protochloride — Mild Chloride—Calomel — Hydrargyri Chloridum** (U.S.)—**Hydrargyri Subchloridum** (Br.), ($HgCl$ or Hg_2Cl_2), is usually prepared by the mutual decomposition of sodium chloride, mercuric sulphate and mercury. After mixing thoroughly in a mortar, the mixture is heated, when the calomel sublimes.

$$HgSO_4 + 2NaCl + Hg = Na_2SO_4 + Hg_2Cl_2.$$

By this method, mercuric chloride is also formed in varying quantities, and should be removed by washing the product with boiling distilled water, until the washings no longer form a precipitate with ammonium hydrate.

Mercuric chloride may be detected in calomel by its forming a black stain upon a bright iron surface dipped in a mixture of calomel and alcohol; or, by the production of a black stain by hydrogen sulphide in water that has been filtered through calomel so contaminated. Calomel crystallizes when sublimed, in radiating quadratic prisms; but if precipitated from solutions of mercurous salts by hydrochloric acid, it forms a heavy, white, amorphous powder. Heated to about 500° C. (932° F.), it sublimes without fusing, is insoluble in cold water and alcohol, and dissolves in boiling water to the extent of 1 part in 12,000. If boiled for a long time with water, it partly decomposes, mercury being deposited, and mercuric chloride passing into solution. Strong acids convert it into mercuric salts and free mercury. With ammonium hydrate it blackens, with formation of **mercur-amidogen chloride.**

$$Hg_2Cl_2 + 2NH_4OH = NH_4Cl + NH_2Hg_2Cl + 2H_2O.$$

Hydrochloric acid and alkaline chlorides convert it into mercuric chloride; this may occur in the stomachs of persons who use large quantities of salted food, as on board ship. Alkaline iodides convert it first into mercurous iodide, which is then decomposed by an excess of the alkaline iodides into mercuric iodide and mercury.

477. **Mercurous Iodide—Protoiodide, or Yellow Iodide — Hydrargyri Iodidum Viride** (U. S., Br.), ($Hg\bar{I}$ or Hg_2I_2), is prepared by triturating 200 parts of mercury with 127 parts iodine and a little alcohol, until a green paste is formed. It may also be prepared by precipitation from a solution of $Hg_2(NO_3)_2$ with KI. It is a greenish-yellow, amorphous powder, insoluble in water and alcohol. It turns brown and volatilizes when heated. Light decomposes it into mercuric iodide and mercury.

478. **Mercurous Oxide—Protoxide, or Black Oxide** (Hg_2O), is formed by the action of sodium hydrate upon mercurous salts. It is a brownish-black, tasteless powder, which sunlight decomposes into mercuric oxide and mercury. Mineral acids convert it into the corresponding mercurous salts. It exists in the **lotio hydrargyri nigri** (Br.), or **black wash.**

479. **Mercurous Nitrate** ($HgNO_3$ or $Hg_2(NO_3)_2$) is formed by digesting an excess of mercury with somewhat diluted nitric acid, until short prismatic crystals separate. The crystals effloresce in the air. Water decomposes this salt into the acid salt, which goes into solution, and **basic mercuric nitrate** $\left(Hg\begin{matrix}\diagup OH\\ \diagdown NO_3\end{matrix}\right)$, which separates as a yellow powder. Water acidulated with nitric acid dissolves it, but it soon oxidizes and becomes mercuric nitrate. By adding metallic mercury to the solution, this oxidation is prevented to a great degree, or after oxidation reduces it back to the **ous** state.

$$Hg(NO_3)_2 + Hg = Hg_2(NO_3)_2.$$

480. **Mercurous Sulphate** ($HgSO_4$) is formed by gently warming an excess of mercury with sulphuric acid. It separates as a yellow crystalline precipitate when sulphuric acid is added to mercurous nitrate solution.

481. **Mercuric Compounds.**—In these, mercury is bivalent; they are represented by the formula HgR_2. The mercuric compounds are always formed when mercury is dissolved in excess of acid; when the opposite is the case, the **ous** compounds

[illegible] metallic mercury [illegible]
[illegible] mercury [illegible]

[illegible] Chloride — Bichloride — Corrosive S
[illegible] Chloridum Corrosivum [illegible]
[illegible] $HgCl_2$ [illegible] may be produce
[illegible]
[illegible] by sublimating a dried mi
[illegible] sodium chloride [illegible]

$2NaCl$ [illegible] HgS [illegible] $= Na_2S$ [illegible]

[illegible] rectangular octahedra [illegible] needle-like prisms [illegible] parts of water and [illegible] which [illegible] It appears [illegible] mercury and [illegible] Subl [illegible] change. The sublimate [illegible] without [illegible] white precipitate [illegible]

[illegible] Chloride — White Pre
[illegible] Mercury. Hydrargyrum Amm
[illegible] as a heavy [illegible] ammonium hydrat [illegible] insoluble in alcohol [illegible] decomposes it, with the sepa [illegible] sublimes without fusing.

[illegible] Biniodide, or Red Iodi
[illegible] Rubrum (U. S., Br.). HgI
[illegible] mercury and iodine, wh [illegible] When solutions of mer [illegible] are mixed, a double decompo [illegible] separates as a yellow [illegible] bright red. It is spar [illegible] solutions of KI [illegible] It also dissolves in many [illegible] salts, alkaline chl [illegible] salts. From its alcoholic solution it crysta

in bright red, rhombohedral crystals. When heated, it becomes yellow, fuses, and sublimes in yellow, shining, rhombic needles. These again become red upon touching them with some solid, and are changed into a mass of octahedra. Mercuric iodide is, therefore, **dimorphous.**

485. **Mercuric Oxide—Red, or Binoxide—Hydrargyri Oxidum Flavum** (U. S., Br.)—**Hydrargyri Oxidum Rubrum** (U. S., Br.), (HgO), is obtained by igniting mercurous or mercuric nitrate as long as fumes are given off; or, by adding sodium hydrate to a solution of a mercuric salt. The product obtained by the first method is red and crystalline, of sp. gr. 11.2; that obtained by precipitation is yellow and amorphous, furnishing **Hydrarg. oxid. flavum** (U. S. P.). The latter is the more active form. Both modifications turn black when exposed to light and air. At 400° C. (752° F.) it breaks up into mercury and oxygen. Mercuric oxide is very sparingly soluble in water. It is the chief ingredient of the **Lotio hydrargyri flava** (Br.), or **yellow wash,** prepared by adding lime water to a solution of mercuric chloride.

486. **Mercuric Nitrate** ($Hg(NO_3)_2$ may be obtained by dissolving mercury or mercuric oxide in hot nitric acid. This should be carefully conducted as it is inclined to form basic salts. It dissolves in water, and exists in the **Liq. hydrargyri nitratis** (U. S.), or **Liq. hydrargyri nitratis acidus** (Br.). It is used in the volumetric estimation of urea by Liebig's method. The standard solution used for this purpose contains 71.48 grms. of metallic mercury to the litre, and 1 c.c. corresponds to 10 mgrms. of urea.

487. **Mercuric Sulphate** ($HgSO_4$) is prepared by warming mercury or its oxide with an excess of sulphuric acid. With an excess of water, it decomposes into sulphuric acid and the yellow, insoluble, basic salt, **Turpeth mineral,** $HgSO_4.2HgO$. It is a white, crystalline salt, used in some forms of galvanic batteries as the exciting agent.

488. **Mercuric Sulphide—Red Sulphide—Cinnabar—Vermilion—Hydrargyri Sulphuretum Rubrum** (U. S.), (HgS), occurs native in radiating or amorphous masses. It may be prepared by rubbing sulphur and mercury together, or by the precipitation of a mercuric salt by hydrogen sulphide, as a black amorphous mass, which is the **Æthiops mineralis** of the older pharmacists.

489. **Medical Action of Mercury.**—If introduced into the animal economy, metallic mercury is not poisonous. By

contact with alkaline chlorides, however, it is converted into mercuric chloride; the more finely divided the particles of mercury are, the more readily does this take place.

Mercuric chloride has a decidedly toxic action, both locally and constitutionally. Its local irritant action is due to its tendency to unite with albuminoid bodies. The constitutional symptoms are somewhat similar to those produced by arsenic, but appear sooner. The vomit frequently contains blood, and there is an intense burning, metallic taste in the mouth. The symptoms that are referable to the gastro-intestinal mucous membrane are more intense. The size of the minimum fatal dose of the corrosive chloride is about three grains; of white precipitate, thirty to forty grains; and of Turpeth mineral, about forty grains. Children tolerate mercury much better, in proportion to their age, than adults.

The treatment should consist in the administration of milk or white of egg, and the induction of prompt emesis. Absorbed mercury probably exists in the blood as an albuminate, and is eliminated by the fæces, urine and saliva; chiefly by the former. Chronic mercurial poisoning, known as **mercurial tremors, shaking palsy**, etc., is met with in those who work in mercury compounds. The symptoms usually begin with debility, nausea, vomiting, colicky pains, and a constant metallic taste in the mouth. Sooner or later salivation will become a prominent symptom, the tongue and gums becoming swollen, red and ulcerated, and the breath will emit a peculiar fetid odor. Salivation may, however, be produced by bromine, antimony, lead, prussic acid, etc.

Chronic and even acute poisoning may occur from the free external use of mercuric salts. **Post-mortem** the mucous membrane of the stomach is usually found of a grayish color, as also that of parts of the mouth and œsophagus. The surface of the membrane is sometimes covered with a slate-colored deposit of finely divided mercury.

Tests.—One of the simplest tests for mercury in solution is a piece of bright copper, which, in the presence of a small quantity of free hydrochloric acid, becomes coated with a silver white layer of copper amalgam. All salts of mercury are volatile. When heated in a tube with sodium carbonate, globules of metallic mercury distill off from all salts. Mercury salts give a black precipitate with H_2S, which is insoluble in nitric acid, but soluble in aqua regia.

SILVER.

Argentum, Ag = 108.

490. **Occurrence.**—This metal occurs in nature in combination with chlorine, bromine, iodine, sulphur, arsenic, copper, antimony, etc. The principal localities in which it is found are the western United States, Mexico, Hungary and Saxony.

491. **Preparation.**—For an elaborate description of the process by which silver is extracted from its ores, the student is referred to works on metallurgy. As usually obtained by these processes the metal is not pure, but is contaminated to a greater or less extent by copper and other metals. To obtain it chemically pure, the ordinarily occurring silver is dissolved in nitric acid, and from this solution of the nitrates, silver is precipitated as chloride by hydrochloric acid or common salt. The silver chloride thus obtained may be reduced by fusion with sodium carbonate, or by the action of zinc or iron in the presence of water.

$$2AgCl + Zn = ZnCl_2 + 2Ag.$$

492. **Properties.**—A brilliant white metal; sp. gr. 10.47 to 10.54. It is tolerably malleable, soft, very ductile, and is the best conductor of heat and electricity known. It does not oxidize in the air, but frequently tarnishes in ordinary atmospheres, from the presence of minute quantities of hydrogen sulphide, which blackens it. With the members of the **halogen group** it unites directly. It dissolves in hot, strong sulphuric acid, to form the sulphate, but is more easily attacked by nitric acid, which dissolves it with great readiness, even when largely diluted. In order to give it the necessary hardness for use in the arts, it is usually alloyed with copper. Coin silver contains 10 per cent. of copper.

493. **Silver Oxide—Silver Monoxide—Silver Protoxide—Argenti Oxidum** (U. S., Br.), (Ag_2O), is precipitated from solutions of soluble silver salts by sodium or potassium hydrate, as a dark brown, faintly alkaline powder, slightly soluble in water. It has strong basic properties. It readily gives up its oxygen when heated. When it is dissolved in ammonium hydrate there separate, on evaporating, black crystals of an explosive compound ($Ag_2O.2NH_3$), which, when dry, explode on the slightest disturbance.

Silver Suboxide (Ag_4O) and **Silver Peroxide** (AgO, or Ag_2O_2) are also known.

form. The addition of metallic mercury to the mercuric compounds converts them into mercurous compounds, while oxidizing agents produce the opposite effect.

482. **Mercuric Chloride—Bichloride—Corrosive Sublimate—Hydrargyri Chloridum Corrosivum** (U. S.), **Hydrargyri Perchloridum** (Br.), ($HgCl_2$), may be produced by dissolving mercuric oxide in hydrochloric acid.

On a large scale, it is prepared by subliming a dried mixture of mercuric sulphate and sodium chloride.

$$HgSO_4 + 2NaCl = HgCl_3 + Na_2SO_4.$$

It crystallizes by sublimation in rectangular octahedra; from solution in fine, right rhombic, needle-like prisms. At ordinary temperatures, it dissolves in 15 parts of water, and at 100° C. (212° F.), in 3 parts; it is still more soluble in alcohol. It dissolves freely in hot hydrochloric acid, which solution gelatinizes on cooling. Its specific gravity is 5.4. In aqueous solution, it tends to reduce into metallic mercury and calomel. Sodium or ammonium chloride prevents this change. Zinc, cadmium, nickel, iron, lead, copper and bismuth, remove most of its chlorine, reducing it either to metallic mercury or calomel. Sulphuric, nitric and hydrochloric acids all dissolve it without decomposition. With albumen, it forms a white precipitate, insoluble in water, but soluble in excess of albumen solution, or in solutions of alkaline chlorides.

483. **Mercur-Ammonium Chloride—White Precipitate—Ammoniated Mercury—Hydrargyrum Ammoniatum** (U. S., Br.), (NH_2HgCl), is thrown down as a heavy white precipitate, by adding a slight excess of ammonium hydrate to a solution of mercuric chloride. It is insoluble in alcohol, ether, and cold water. Hot water decomposes it, with the separation of a heavy yellow powder. It sublimes without fusing.

484. **Mercuric Iodide—Biniodide, or Red Iodide—Hydrargyri Iodidum Rubrum** (U. S., Br.), (HgI_2), is formed by the direct union of mercury and iodine, when the two are triturated in a mortar. When solutions of mercuric chloride and potassium iodide are mixed, a double decomposition takes place, and the mercuric iodide separates as a yellow precipitate, which immediately turns bright red. It is sparingly soluble in water, but freely soluble in solutions of KI and alcohol, forming clear solutions. It also dissolves in many dilute acids, and in solutions of ammonium salts, alkaline chlorides and mercuric salts. From its alcoholic solution it crystallizes

floating over it a solution of collodion (a solution of **pyroxyline** in ether and alcohol) containing a small quantity of iodide or bromide of potassium, and then dried. On dipping the plate into a solution of silver nitrate, it becomes coated with a layer of silver iodide or bromide. After exposure in the camera it is taken to the dark room and "**developed**" by pouring upon it a solution of pyrogallic acid, or ferrous sulphate, which reduces the silver salts to the metallic state on that portion of the plate that has been acted upon by light, and makes it opaque. In printing from this, a paper previously wetted by floating it upon a solution of silver nitrate, is placed behind the negative and exposed to a strong light. The same action takes place upon the paper; the lights upon the negative becoming dark upon the paper. The image is fixed by dissolving off the undecomposed silver with a solution of sodium hyposulphite.

GOLD.

Au = 197.

500. This is the only member of the group. It is sometimes classified as one of the precious metals.

501. **Occurrence.**—Gold occurs native in wide distribution, though in small quantities. Generally found in veins of quartz, from which it is extracted, after pulverization, by dissolving it out with mercury. The amalgam thus formed is subjected to heavy pressure, which squeezes out the excess of mercury, the rest being separated by distillation.

502. **Properties.**—Gold is orange yellow by reflected light, and green by transmitted light; very ductile and exceedingly malleable. It fuses at 1200° C. (2192° F.); has a specific gravity of 19.36, and is a good conductor of heat and electricity. It retains its lustre even at high temperatures. It is not affected by any single acid or alkaline hydrate. A mixture of nitric and hydrochloric acids readily dissolves it, forming a solution of the chloride. It combines directly with the halogens, phosphorus, antimony, arsenic and mercury. In handling bromine, care should be taken that its vapor, or the bromine itself, does not come in contact with rings or other gold jewelry, lest they be attacked.

503. **Uses.**—Neither gold nor its preparations are much used in medicine. It is extensively employed, however, in the manufacture of jewelry and for coinage. For either of these purposes it is too soft to be used alone, but is always alloyed with either copper or silver. In estimating the fineness of gold in jewelry,

[illegible]

[illegible] carats. [illegible]
carats [illegible]
metal [illegible]
and six carats [illegible]
ark violet [illegible]
n aurous chloride. [illegible]
le and gold.
pared by digesting [illegible]
decomposing the [illegible]
residue at 100° C. 212° F.
decomposes easily, and [illegible]
m **aurates** having the [illegible]
be seen, from its behavior
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is a pale yellow, insoluble
oride to 200° C. (392° F.)
C. NaCl is used in medicine.
aving a saline, metallic taste
water and alcohol
Trichloride — AuCl, occurs
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readily decomposed, with
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nt precipitate—**purple of**
glass and porcelain.

GROUP.

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113.4

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aminium may be
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better, the double chloride of sodium and aluminium with sodium.

$$Al_2Cl_6.2NaCl + 3Na = Al_2 + 8NaCl.$$

It is a silver-white metal, very malleable and ductile, a good conductor of electricity; sp. gr. 2.56. At ordinary temperatures it is not affected by air or oxygen, but burns if heated in oxygen. It is insoluble in nitric acid, soluble in boiling sulphuric and cold hydrochloric acids. It dissolves in alkaline hydrates to form **aluminates,** with liberation of hydrogen. It forms a very hard and durable alloy with copper, known as **aluminium bronze.**

510. **Aluminium Oxide—Aluminia** (Al_2O_3), is found crystallized in prisms, and colored by other admixtures, in **ruby, sapphire, corundum** and **topaz.** These minerals are all exceedingly hard, ranking next to the diamond in this respect. Aluminia may be obtained artificially by igniting the hydrate, as a light, white, insoluble, odorless and tasteless powder. Acids attack it with great difficulty when prepared as above. It may be decomposed by fusing with caustic alkalies or acid potassium sulphate.

511. **Aluminium Hydrate** ($Al_2(OH)_6$) is formed by precipitating a solution of an aluminium salt with ammonium hydrate or carbonate. When freshly precipitated it is insoluble in water, but soluble in solutions of the fixed alkalies. By prolonged drying, or after standing under water, it is rendered almost insoluble in acids, although it undergoes no change in composition or appearance.

512. **Aluminium Chloride** (Al_2Cl_6) is obtained by the action of chlorine upon heated aluminium. It forms colorless hexagonal prisms; fusible, volatile, and very soluble in water and alcohol. It crystallizes from a hot concentrated solution with $12H_2O$.

513. **Aluminium Sulphate** ($Al_2(SO_4)_3$) is prepared artificially by the action of sulphuric acid upon kaolin or clay. Also, by dissolving aluminium hydrate in the same acid (aluminii sulphas, U. S.). It crystallizes in thin plates with $18H_2O$; soluble in water, and sparingly so in alcohol. When heated, it fuses and becomes anhydrous.

514. **Alums.**—These are double salts formed by the combination of aluminium sulphate with the alkaline sulphates. The salt originally known as **alum** is the double sulphate of aluminium and potassium, ($K_2Al_2(SO_4)_4.24H_2O$). It is obtained from clays free from iron, and from **aluminite,** a basic sulphate of

. . . be replaced by thallium. Potas- . . . tahedral crys- . . . it loses 45 . . . forming . . . known as . . . **and ammonium sulphate,** . . . NH_4 $SO_4.24H_2O$, is rapidly . . . alum, from which it differs in . . . 20°C. (68°F.) and 90°C. . . . cooler or warmer than this. . . . tash alum loses its water, . . . its ammonia. Ferric iron, . . . place the aluminium in alum, . . . as **ferric alum** (see Art. . . . **hrome alum.**

NUM METALS.

Rhodium, Rh = 104.
Ruthinium, Ru = 104.
197.2

. . . ually classed as "metals of the . . . **group of metals.** The . . . ars in small metallic grains in the Ural mountains, Brazil . . . hydrates (or salts represent- . . . and therefore play both the . . . is the only one used to

UM.

. . . etc.—It exists in nature . . . group; also, in ores con-

. . . gr. 21.5. It is very . . . able of being drawn out . . . res it softens without . . . Upon heating the . . . um, a grayish-black, . . . **nge.** This latter has . . . of certain gases. A . . . mes by the oxidation of the hydrogen in the pores of the sponge, the heat developed being sufficient to inflame the hydrogen. Platinum is not affected by the air or oxygen. It unites with chlorine, arsenic, silicon, sulphur and phosphorus. It does not dissolve in acids, and is only soluble in liquids generating free chlorine, as aqua regia. With many heavy metals it forms easily fusible alloys. Therefore, easily reducible metallic oxides, as of arsenic, lead, etc., ought never to be ignited in platinum vessels. Platinum is valuable for its high fusing point and its power to withstand oxidation; and is expensive, because of its scarcity.

517. **Platinum Compounds.**—These are of two series, platinous, PtR_2, and platinic, PtR_4. In the first, the metal is more basic, and in the latter, more acid in nature.

Platinic Tetrachloride ($PtCl_4$) is obtained by dissolving platinum in aqua regia and evaporating off the nitric acid. On evaporating, it crystallizes in soluble, reddish-yellow needles. With ammonium or potassium chloride and the alkaloids, it forms characteristic double chlorides. This salt is largely used to precipitate potassium, ammonium, or the alkaloids, for quantitative estimation.

MOLYBDENUM GROUP.

Molybdenum, Mo = 96. **Tungsten** (Wolfram), **W = 184.**
Osmium, Os = 198.

518. These elements are of little importance themselves; their oxides, however, and some of their compounds are sometimes used.

519. **Molybdic Trioxide or Anhydride** (MoO_3).—This oxide is obtained by roasting the native sulphide in an open vessel, at a red heat. The principal interest attached to it is its use in preparing ammonium molybdate, a reagent used to detect and estimate phosphoric acid.

The impure oxide obtained by roasting the mineral **molybdenite**, MoS_2, is dissolved in ammonium hydrate, evaporated to dryness, re-dissolved in water, concentrated by evaporation, and allowed to crystallize. The soluble molybdates give a precipitate of MoO_3 on the addition of acid; but it is soluble in excess of acid. The molybdates give a white precipitate with the earthy metals. With phosphoric acid or the phosphates, a solution of ammonium molybdate containing an excess of nitric or hydrochloric acid first turns yellow, then deposits a yellow precipitate of molybdic trioxide, phosphoric acid and ammonia, [illegible] soluble in ammonia water. This is a very delicate [illegible] pyrophosphates and metaphosphates [illegible] in solution. Arsenic acid gives a similar precipitate.

[illegible] Tungstic acid [illegible] is a yellowish-white powder [illegible] of tungstic oxide [illegible] It forms with bases numerous salts, called tungstates, the most important of which is sodium tungstate [illegible]. This has recently attained considerable importance [illegible] and from its property [illegible] non-inflammable.

OSMIUM.

Os – 195

[illegible] This unusual metal, occurring with some of the other platinum metals, is mentioned here for the sake of its oxide [illegible], which has received considerable attention as a staining agent. [illegible]

The tetroxide [illegible], so-called osmic acid, may be prepared by heating the metal in air or by the action of chlorine and moisture upon the metal. This oxide has a very irritating, piercing odor, and is extremely poisonous. It has been used to arrest the motions of and kill the organisms in water, for microscopical examination.

Many kinds of organic substances reduce this oxide and precipitate the metal, and it is due to this property that it has found use as a staining agent in histology, to bring out delicate structures.

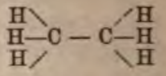
PART IV.

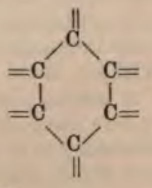
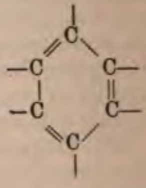
ORGANIC CHEMISTRY.

522. **Carbon and Hydrogen.**—The substances derived from the animal and vegetable kingdoms are composed principally of four elements: carbon, hydrogen, oxygen and nitrogen, with occasionally sulphur and phosphorus. Few as are the chemical elements concerned, the number of different compounds of these elements is almost endless. Formerly only those substances directly or indirectly derived from bodies possessing vegetable or animal life were considered as organic; but as our knowledge increased, a large number of compounds were prepared in the laboratory from these bodies, which were identical with others prepared from the elements themselves. It was thus demonstrated that by pure chemical agencies many of the products of living organisms could be prepared, and that a vital principle was not necessary to form these compounds; and, also, that by the same chemical action a great many compounds could be formed which could not be found in either animal or vegetable organisms. Some artificial bodies were afterwards found to exist in either animal or vegetable organisms. This greatly enlarged the field of investigation, and a new meaning was attached to the term **organic chemistry**; it was found necessary to extend the science to include all bodies in any way resembling organic compounds, either in composition or properties; *i. e.*, all compounds containing carbon and hydrogen. Carbon and hydrogen are the indispensable elements to the formation of organic bodies; without these we can have no substance capable of vegetable or animal life. In the great majority of organic compounds we also have oxygen or nitrogen, or both.

523. **Organic Chemistry** may be defined as that branch of the science of chemistry which treats of the carbon compounds containing hydrogen, either alone, or with oxygen, nitrogen, sulphur, phosphorus, etc.

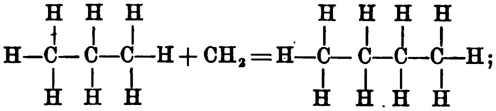
While inorganic compounds, as a rule, contain but a few atoms, organic compounds frequently contain a large number; and the diversity in organic chemistry is obtained not by varying the kind of atoms, but by varying the arrangement of a few kinds. There exist, however, certain organized bodies which possess a structure entirely different from that of any known inorganic body or artificial organic substance. These *organized* structures are the sole product of vital action, and cannot be produced in the laboratory. Such bodies are seen in the cell of living organisms; and although we may be able at some time to construct molecules identical with those composing them, we shall never be able to impart that function of growth, reproduction, and other vital processes which we call life. The complexity of organic compounds may readily be attributed to the inherent properties of carbon, the leading or most characteristic element of which they are composed. The atoms of carbon exhibit, in a remarkable degree, a tendency to combine themselves into groups or chains, around which the other atoms or groups of atoms are arranged. In other words, the carbon atoms are the skeleton in and upon which a very complex structure may be built. The underlying principle of organic chemistry is the grouping of the carbon atoms, and upon these groups the whole superstructure is based.

524. **Carbon Nuclei.**—Carbon is a tetrad and can combine with four monads. As has been said, it can also unite with another carbon atom, and the two atoms thus united will combine with six monad or hydrogen atoms: 

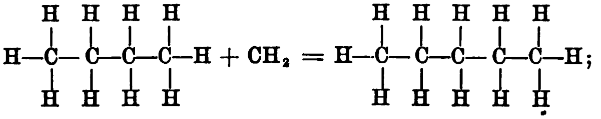
Three, four, five, or any number (n) atoms of carbon may thus unite, theoretically at least, and give rise to a chain of atoms, which has been called an **open chain.** Again, the same atoms may combine as follows:  or 

giving rise to **closed chains.**

By a little consideration it will be seen that, in either of the first two formulæ given above, if we increase the carbon atoms in any given formula, we also increase the number of **bonds,** or points of contact for hydrogen atoms, by two for every carbon atom added. Thus:



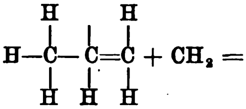
$$H{-}CH_2{-}CH_2{-}CH_2{-}H + CH_2 = H{-}CH_2{-}CH_2{-}CH_2{-}CH_2{-}H;$$



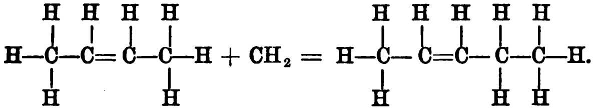
$$H{-}CH_2{-}CH_2{-}CH_2{-}CH_2{-}H + CH_2 = H{-}CH_2{-}CH_2{-}CH_2{-}CH_2{-}CH_2{-}H;$$

and so on.

We thus have a series of hydrocarbons differing from each other by the constant quantity of CH_2. Such a series is called a **homologous series.** In the same chain of carbon atoms, we might have the following arrangement of the bonds of attraction:—



$$H{-}CH_2{-}CH{=}CH_2 + CH_2 =$$



$$H{-}CH_2{-}CH{=}CH{-}CH_2{-}H + CH_2 = H{-}CH_2{-}CH{=}CH{-}CH_2{-}CH_2{-}H.$$

This gives rise to another homologous series, each member differing from the last mentioned by the constant quantity of H_2, and known as the **isologous series.** In the vertical columns of the following table we have the **homologous series**; in reading from left to right, we have the **isologous series** of hydrocarbons.

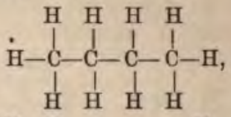
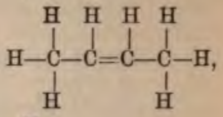
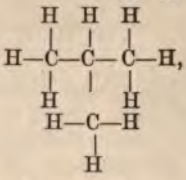
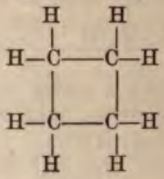
CH_4	CH_2*	C*	
C_2H_6	C_2H_4	C_2H_2	
C_3H_8			
C_4H_{10}	C_3H_6	C_3H_4	C_3H_2
C_5H_{12}	C_4H_8	C_4H_6	C_4H_4
C_6H_{14}	C_5H_{10}	C_5H_8	C_5H_6
C_7H_{16}	C_6H_{12}	C_6H_{10}	C_6H_8
C_8H_{18}	C_7H_{14}	C_7H_{12}	C_7H_{10}
C_nH_{2n+2}	C_nH_{2n}	C_nH_{2n-2}	C_nH_{2n-4}

* Unknown.

This table does not by any means represent all the possible compounds, but gives an idea of how all known hydrocarbons may be classified by extending the table.

525. **Isomerism, Polymerism.**—Besides the method of the carbon nuclei which we have indicated, a variety of other

methods may be conceived, some of which really exist. Thus the group containing four carbon atoms may be arranged as follows:

The first and third of these would have the same number of atoms and the same formula, C_4H_{10}, but would have different physical and chemical properties. Compounds having the same chemical composition but possessing different properties, are termed **isomeric bodies.** The same may be said of the second and fourth of the above formulæ; the common formula is C_4H_8, but they are quite different in the constitution of their molecules, while they are the same in composition. These two compounds are isomeric. The chemical and physical properties of organic compounds depend not only upon the kind and number of atoms, which is designated the **composition** of the molecules, but also upon the arrangement of the atoms in their molecules, or their **constitution.** It is in this way that we may have several compounds, each answering to the same formula, and giving the same result on analysis, but totally differing in properties. We may mention here, as examples, ammonium cyanate, NH_4CyO, and urea, $CO\left\{\begin{matrix}NH_2\\NH_2\end{matrix}\right.$.

Also,

$$N\left\{\begin{matrix}C_3H_7\\H\\H\end{matrix}\right. \qquad N\left\{\begin{matrix}CH_3\\C_2H_5\\H\end{matrix}\right. \qquad N\left\{\begin{matrix}CH_3\\CH_3\\CH_3\end{matrix}\right.$$

Propylamine Methyl-ethylamine Trimethylamine.

Polymerism is a term given to compounds having the same percentage composition but different molecular weights. We have a striking example of this property in the second column of the table above. Ethene, C_2H_4, Butene, C_4H_8, and Octene, C_8H_{16}, are polymeric bodies. Starch, cellulose, dextrine (British gum), and glycogen of the liver are polymeric bodies.

526. Nomenclature of Homologous Hydrocarbons.—The following are the names proposed for these hydrocarbons:

1st Series.	*2d Series.*	*3d Series.*	*4th Series.*	
PARAFFINES.	OLEFINES.	ACETYLENES.		
C_nH_{2n-2}	C_nH_{2n}	C_nH_{2n-2}	C_nH_{2n-4}	
Methane, CH_4.	Methene* (unknown). CH_2.	C		
Ethane, C_2H_6.	Ethene, C_2H_4.	Ethine (Acetylene), C_2H_2.		
Propane, C_3H_8.	Propene, C_3H_6.	Propine, C_3H_4.	Propone, C_3H_2.	
Quartane[1], or Butane, C_4H_{10}.	Butene, C_4H_8.	Butine, C_4H_6.	Butone, C_4H_4.	Quartune, or Butune, C_4H_2.
Quintane, or Pentane, C_5H_{12}.	Pentene, C_5H_{10}.	Pentine, C_5H_8.	Pentone, C_5H_6.	Pentune, C_5H_4.
Sextane, or Hexane, C_6H_{14}.	Hexene, C_6H_{12}.	Hexine, C_6H_{10}.	Hexone, C_6H_8.	Hexune, C_6H_6.

[1] Butane, pentane, etc., are preferred.

* Also called methylene, ethylene, etc.

Here, as in inorganic chemistry, there is not perfect uniformity in nomenclature, and, where they conflict, we shall follow the best usage in preference to strict conformity to the above.

527. The Paraffines.—The paraffines are chemically indifferent bodies, because they are fully saturated compounds. Compounds derived from these bodies can only be formed by breaking up the molecule or removing one or more atoms and **substituting** other atoms for them. To the medical student, but one member of the group has sufficient interest to be studied in detail: viz., methane, "**marsh gas,**" or **light carburetted hydrogen.** This gas is a constant product of the decomposition of vegetable matter under water, or of the destructive distillation of coal, and hence is a constituent of illuminating gas. It is the chief constituent of miner's "fire damp." It is very combustible in the air, burning with a pale bluish flame, and giving a high temperature but little light. When mixed with air, it forms an explosive mixture. It may be prepared by heating a mixture of 4 parts sodium acetate, 4 parts sodium hydrate and 6 parts of lime in powder.

$$NaC_2H_3O_2 + NaOH = CH_4 + Na_2CO_3.$$

It is a colorless, nearly odorless gas. Density = 8; sp. gr. = 0.559. It is not poisonous.

The higher members of this series are found in American petroleum, and may be isolated to a great extent by careful fractional distillation. The distillation of petroleum is carried on on a large scale, in certain parts of this country, and a variety of products are found in the market, the names of which do not have any reference to the chemical composition. The following are a few of the most important of these products, with their boiling points and uses:—

Name.	Boiling Point.	Principal Use in Arts.
Cynogene	0° C. (—32° F.)	Used in ice machines.
Rhigolene	18.3° C. (65° F.)	Used to produce cold by evaporation, and as an anæsthetic.
Gasolene	48.8° C. (119° F.)	For making "air gas."
C Naphtha	82.2° C. (180° F.)	As a solvent for fats and rubber, hence for cleaning cloths. So-called "safety oil."
B Naphtha	104.4° C. (220° F.)	
A Naphtha	148.8° C. (300° F.)	
Benzene	148° to 160° C. (298° to 320° F.)	For varnishes and paints.
Kerosene, or Refined Petroleum	176° C. (349° F.)	For ordinary lamps.
Mineral Sperm Oil	218° C. (424° F.)	
Lubricating Oil	301° C. (574° F.)	Lubricating machinery.
Paraffine	Solid.	For candles.

The vapors of all the lighter products when combined with air form explosive mixtures, and hence, laws exist in most countries prescribing the lowest temperature at which kerosene shall give an inflammable vapor, or at which it shall "flash." The law of the State of New York declares that oils used for illuminating purposes shall not give a vapor that will "flash" below 100° F., and shall not themselves ignite below a temperature of 300° F. Fossil resins, bitumen, ozocerite (animal wax), all belong to this group of hydrocarbons.

528. **The Olefines, or C_nH_{2n} Series.**—The olefines differ from the paraffines in being unsaturated compounds. The first member of the series, **ethylene**, C_2H_4, combines directly with chlorine, forming a thick, oily liquid, from which the discoverers named it olefiant (oil making) gas. The iodide and bromide may also be formed by direct union. The olefines combine with the hydracids.

$$C_2H_4 + HCl = C_2H_5Cl.$$

Ethylene (olefiant gas), C_2H_4, is found in illuminating gas, the illuminating power of which depends largely upon its presence. It is a colorless gas, of a peculiar, pungent odor, and is soluble in strong H_2SO_4, with which it combines. It burns in the air with a bright, luminous flame. Its sp. gr. is .9785; density, 14.

The higher members are unimportant; of these butylene, C_4H_8, is liquid or solid at ordinary temperatures.

529. **Acetylenes, C_nH_{2n-2} Series.**—This series of hydrocarbons fall short of saturation by four monad atoms, and can, therefore, act as dyad or tetrad radicals. Acetylene, C_2H_2, combines directly with either two or four atoms of chlorine, bromine or iodine. Nascent hydrogen converts it into ethylene, C_2H_4, and oxidizing agents (potassium permanganate, etc.) into oxalic acid. Acetylene is formed whenever kerosene or coal gas is incompletely burned, and gives the disagreeable odor detected in a room when the lamp is turned low. It exists in coal gas, and possesses a high illuminating power.

530. **Fourth Series.**—General Formula, C_nH_{2n-4}. This series of hydrocarbons includes turpentine and a large number of other so-called essential, or volatile oils. These oils are mostly isomers or polymers, having the formula $C_{10}H_{16}$, or a multiple of this.

Turpentine is extracted from several varieties of the conifera family, notably the pine. When turpentine is distilled, the hydrocarbons volatilize, while resin remains behind. Oil of turpentine is a mobile, colorless liquid, with a sp. gr. of 0.89, and boils at 160° C. (320° F.). It is almost insoluble in water, but dissolves in alcohol, ether, and glacial acetic acid. It dissolves sulphur, resin and phosphorus. It absorbs oxygen from the air and becomes oxidized, forming a resinous body. The absorbed oxygen is converted into ozone, and this explains its oxidizing, disinfectant and antiseptic action. As the oxidation of turpentine takes place more rapidly when mixed with lead oxide, this oxide is often intermixed with the turpentine in paints, to increase the rapidity of drying. It attacks lead rapidly, but not tin.

A paper dipped in turpentine and introduced into a jar of chlorine gas inflames spontaneously, forming substitution products. Iodine and bromine have a similar action upon it. It unites directly with HCl, producing several hydrochlorides.

Oil of Lemon, obtained from the peel of the common lemon, *Citrus Limonum*, oil of orange peel, cloves, juniper, pepper, elemi, lavender and bergamot, belong to the same class, having the formula $C_{10}H_{16}$. All the essential oils, however, do not belong to this series of hydrocarbons. Oils of copaiva and cubebs have the formula $C_{20}H_{32}$, while caoutchouc and gutta-percha are higher polymers of the above.

The principal characteristics of these essential oils are their odors, variations in the rapidity of oxidation, reaction with hydrochloric acid, and physical properties. They are soluble in alcohol, ether, benzole, petroleum, naphtha, chloroform, carbon pisulphide, paraffine, and other volatile oils, and in the fixed oils. They may be separated from their alcoholic solution by adding water, or a solution of sodium sulphate. They are not saponified by alkalies.

531. **Caoutchouc, or India Rubber,** is the dried milky juice of several tropical trees. The fresh juice is acid. It is a mixture of several hydrocarbons which are insoluble in alcohol and water, but soluble in sulphuric ether, benzole, chloroform, carbon disulphide and turpentine. When cold, it is hard and tough, but on heating, it becomes soft, elastic, and finally melts, and on cooling, remains soft and viscid. It is much used in making elastic, water-proof fabrics, elastic tubing, etc., and is acted upon by but few reagents. The black color of the commercial article is due to smoke and partial decomposition.

Caoutchouc combines with sulphur. **Vulcanized** India-rubber is obtained by mixing it intimately with sulphur, by the aid of carbon disulphide, to the extent of two or three per cent., and afterward heating. Common white rubber goods, as rubber tubing, etc., are also mixed with oxide of zinc and other impurities to a very large extent, in some cases but a few per cent. of rubber being used. When mixed with about half its weight of sulphur, a hard, horny mass, called **Vulcanite,** or **Ebonite,** is produced, which is used in the manufacture of combs, cheap jewelry, etc. When heated, caoutchouc decomposes, but does not volatilize.

Gutta-percha is the hardened, milky juice of *Isonandra Gutta*, a tree growing in some parts of India. It resembles caoutchouc, but is harder and less elastic. In hot water, it becomes quite soft and can be moulded into any shape, which it retains on cooling. With solvents and high temperatures, it behaves like caoutchouc.

CAMPHORS.

532. **Common Camphor,** $C_{10}H_{16}O$, is obtained in China and Japan by distilling the branches and leaves of *Laurus Camphora* with water. It is a white, translucent, crystalline mass, having a powerful, peculiar, pungent odor and taste. It is readily

purified by sublimation at 205° C. (401° F.). It melts at 175° C. (347° F.) and burns with a smoky flame. Camphor is very slightly soluble in water, but readily soluble in alcohol, ether, acetic acid, benzole, chloroform, carbon disulphide, fixed and essential oils. **Aqua Camphoræ** and **Tinctura Camphoræ** are officinal, and it enters into the composition of **Linimentum Camphoræ, Liniment. Saponis, Tinctura Opii Camphorata.**

Camphor forms a large number of decomposition products and derivatives under the action of reagents, but we shall notice but one of these.

Monobromo-camphor—Camphora Monobromata (U. S. P.), $C_{10}H_{15}BrO$, is prepared by adding bromine to a solution of camphor in chloroform, by which camphor dibromide is obtained. This compound is unstable, and on standing, sets free hydrobromic acid and forms monobromated camphor, which crystallizes in colorless, prismatic needles or scales, permanent in the air, having a mild camphoraceous odor and taste, and neutral reaction. Its solvents are essentially the same as those of camphor. It melts at 65° C. (149° F.), boils at 274° C. (525° F.), and is volatilized with partial decomposition. In medicine it is used as a sedative.

Eucalyptol, $C_{12}H_{20}O$, is a colorless liquid, boiling at 175° C. (347° F.), and possessing an aromatic odor. It is contained in the leaves of the *Eucalyptus Globulus*, a tree growing in Tasmania. On account of its supposed effect upon miasmatic atmospheres, it has been cultivated in southern Europe, the United States and northern Africa. By distilling it with phosphorus pentachloride, PCl_5, **eucalyptine**, $C_{12}H_{18}$ (by some $C_{10}H_{16}$), is obtained. Eucalyptol is slightly soluble in water, but soluble in alcohol. The oil has feeble antiseptic properties and has been used in bronchitis, cystitis, and in intermittent fevers.

Menthol,* or Menthyl Alcohol, $C_{10}H_{20}O$, is a white, solid, crystalline body occurring in oil of peppermint, and possessing a strong odor of this plant. It melts at 36° C. (96.8° F.) and boils at 210° C. (410° F.) Menthol is soluble in alcohol and the essential oils.

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* Although really an alcohol, it is classed here with the camphors, owing to physical properties.

RESINS, OLEO-RESINS, GUMS, BALSAMS.

533. Many of the bodies of the turpentine and essential oil series above mentioned, when exposed to the air, undergo a process of oxidation or hardening, become viscid or solid, and exhibit an acid reaction.

Such bodies, when brittle and solid, are called **resins** ; when composed of unoxidized oils mixed with resins, they are called **balsams** ; when the resins exist in the juices of plants mixed with gum, sugar, etc., they are called **gum-resins.** Each one of these bodies is generally a mixture of several bodies, and, therefore, no definite chemical formula can be given. They are insoluble in water, but soluble in strong alcohol, turpentine and glycerine ; many are soluble in ether and benzole (separation from gums) ; many are weak acids, whose alkaline salts form the **resin soaps** of the market. Many resins are used in medicine, in the manufacture of varnishes, sealing wax and salves.

The resins soften when heated, but do not vaporize. The separation of resins from volatile oils and acids, is effected by **distillation** with water ; from gums, by **fusion** and straining at 100°C. (212°F.) ; and from each other, as well as from foreign substances, by properly selected **solvents.**

Only the most important resins can be described. **Common resin**, or colophony, is the residue left by distilling the balsam of the pine with water. Turpentine passes over, and leaves a brown, brittle, shining mass, which, when melted, forms common resin. It is soluble in alcohol, ether and oils.

Copal is a yellowish or brown resin, obtained from various trees of the East and West Indies and Africa, and is used in preparing varnishes.

Amber is a yellowish, hard, brittle, more or less transparent solid, found on the shores of the Baltic, and called **fossil resin.**

Ammoniacum contains 72 per cent. resin, 22 per cent. gum, and a little volatile oil.

Benzoin contains 75 per cent. resin, 10 to 15 per cent. of benzoic acid, a little gum and volatile oil.

Balsam Copaibæ consists of several resins and a volatile oil.

Guaiacum is a brittle, pulverizable solid, of a reddish-brown color. The gum dissolves in alcohol. It readily undergoes oxidation, producing bright colors. A mixture of the officinal tincture and oil of turpentine is frequently employed as a reagent for detecting blood in urine, with which it strikes a blue color

Gum Lac, or Shell Lac, is an exudate from several East Indian trees, caused by punctures of insects. It contains about 90 per cent resin, and a coloring matter soluble in water. It is soluble in alcohol, this solution being used as a varnish.

Among the balsams, **storax, tolu,** and **balsam of Peru** may be mentioned. Among the gum resins, the following are most important: **Ammoniac** (see above), **galbanum, gamboge, myrrh** and **asafœtida.**

BENZOL OR AROMATIC SERIES.

C_nH_{2n-6}.

534. One member of this series is of considerable importance: viz., **Benzol,** or **benzene.** This compound, C_6H_6, must not be confounded with **benzene,** which is a trade name, and is used to designate a mixture of hydrocarbons having a variable composition and boiling point. Benzol was formerly prepared by distilling benzoic acid with lime. At present it is obtained in large quantities by the distillation of coal tar from gas works. Benzol is a colorless, limpid, highly refractive liquid, with a peculiar odor and burning taste. It burns with a smoky flame, and is insoluble in water, but soluble in alcohol, ether and naphtha. It dissolves sulphur, phosphorus, iodine, resins and fats.

Benzol, with its homologue, Toluene C_7H_8, is very interesting to the chemist, because they serve as the starting point in the synthesis of a large number of coloring matters known as aniline colors.

NAPHTHALENE AND ANTHRACENE.

535. **Naphthalene,** $C_{10}H_8$, is obtained from coal tar by distillation between 180°C. and 220°C. It comes over with heavy oils, but crystallizes from them in white, glistening, leafy crystals, of a peculiar, aromatic odor and burning taste; it melts at 79.2°C. (174.5°F.), is insoluble in water, but soluble in hot alcohol, ether or benzene.

The principal interest of naphthalene to the physician is its value as an antiseptic dressing for wounds. For this purpose it must be thoroughly purified by recrystallization from alcohol, or by distillation with steam.

Anthracene, $C_{14}H_{10}$, is a white, crystalline body obtained from coal tar, distilling above 360°C. (680°F.).

It is of great interest to the chemist, being the starting point in the manufacture of **alizarine,** or artificial madder. Its solutions possess a beautiful blue fluorescence, a property observed in many of the heavier hydrocarbons derived from coal tar and petroleum; *i. e.*, their solutions are colorless or yellowish by transmitted light, but when viewed by reflected light, appear bluish. The phenomenon is well seen in solutions of quinine sulphate.

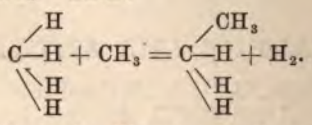
536. **Addition Products** are readily formed from unsaturated hydrocarbons like the olefines or acetylenes. One molecule of the hydrocarbon always unites with two or four atoms or radicals added. It is especially the atoms of the chlorine group, or their hydracids, which are thus added; but hydrogen, as well as other hydrocarbon radicals, may also be added.

$$\underset{\text{Ethylene.}}{C_2H_4} + Cl_2 = \underset{\text{Ethylene chloride.}}{C_2H_4Cl_2.}$$

$$C_2H_4 + HCl = \underset{\text{Ethyl chloride.}}{C_2H_5Cl.}$$

$$C_2H_4 + HOCl = \underset{\text{Ethylene chlorohydrine.}}{C_2H_4\begin{cases}OH\\Cl.\end{cases}}$$

537. **Substitution Products.**—One or more of the hydrogen atoms of the various hydrocarbons may be made to exchange places with either simple or compound radicals. Indeed, we might regard all the complex organic bodies as made up in this way. Thus, ethane, C_2H_6, may be regarded as a molecule of methane, CH_4, in which one hydrogen atom has been replaced with the radical CH_3. Thus:



The limits of this work will not permit us to enter into the various reactions by which all these substitutions are brought about; the student is referred to special works on organic chemistry for these reactions.

538. **Organic Radicals.**—In organic chemistry, as in inorganic, although to a greater extent, we are constantly dealing with certain well defined groups of atoms, which retain their identity through a large series of compounds and behave in chemical reactions like simple radicals or atoms.

It is evident that by removing one or more atoms of hydrogen from any saturated hydrocarbon, the remaining group of atoms may act as a compound radical; while the unsaturated hydrocarbons of the acetylene or olefine series may act as radicals without the removal of any hydrogen atoms.

These hydrocarbon radicals are usually designated by the termination **yl**. Thus: CH_3 is known as methyl; C_2H_5, as ethyl; C_6H_5, as phenyl. These radicals are called, collectively, hydrocarbon or alcohol radicals, and the quality of their combining power is feebly electro-positive.

539. **Synthesis of Organic Compounds.**—The principal interest of chemists in recent years in the department of organic chemistry, is centred upon the synthesis of organic compounds. By these synthetical methods the chemist imitates, to a limited extent, the processes of animal and vegetable organisms. The natural tendency of all organic compounds is to break up complex molecules and form simpler ones, while it is the aim of synthetical methods to reverse this process and build up complex molecules from simpler radicals, adding little by little until the structure is complete.

It is principally from synthetical reactions that we can arrive at a knowledge of the structure of complex organic bodies. With these synthetic methods, chemists have achieved great success in the artificial production of alizarine (the coloring matter of madder), indigo, salicylic acid and the great variety of aniline colors, which stand among the greatest achievements of modern science.

Although we have not space to dwell upon the subject here, we introduce a few reactions from *Cook's Chemical Philosophy*, which will give the student an idea of the succession of steps by which these ends are attained. From the elements, by passing the vapors of sulphur over heated carbon, we obtain CS_2.

From $H_2S + CS_2$ passed over heated copper, we obtain $CuS + CH_4$. The electric arc passed between carbon poles in an atmosphere of hydrogen yields acetylene, $C_2 + H_2 = C_2H_2$.

C_2H_2 + nascent hydrogen $= C_2H_4$.

$$\underset{\text{Ethylene.}}{C_2H_4} + \underset{\substack{\text{Hydrobromic}\\\text{Acid.}}}{HBr} = \underset{\text{Ethyl bromide.}}{C_2H_5Br.}$$

$$C_2H_5Br + \text{nascent hydrogen} = \underset{\text{Ethane.}}{C_2H_6} + HBr.$$

$$\underset{\text{Ethyl Iodide.}}{2C_2H_5I} + \underset{\text{Zinc Methyl.}}{(CH_3)_2Zn} = \underset{\text{Propane.}}{2C_2H_5CH_3} + \underset{\text{Zinc Iodide.}}{ZnI_2.}$$

$$\underset{\substack{\text{Propyl}\\\text{Iodide.}}}{C_3H_7I} + \underset{\substack{\text{Potassium}\\\text{Hydrate.}}}{KOH} = \underset{\text{Propene.}}{C_3H_6} + KI + H_2O.$$

$$\underset{\text{Ethylene Bromide.}}{C_2H_4Br_2} + 2KOH = \underset{\text{Acetylene.}}{C_2H_2} + 2KBr + 2H_2O.$$

These reactions will suffice to show the methods of converting the hydrocarbons from one homologous series to another, or from a lower to a higher member in the same series.

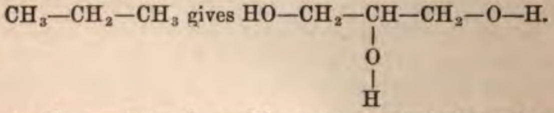
540. **Alcohols.** As shown in Article 319, the hydroxyl group (—OH) may be substituted for one atom of hydrogen in any saturated hydrocarbon. The resulting compound is called an **alcohol.** An alcohol may also be looked upon as one or more molecules of water in which one of the two atoms of hydrogen has been replaced by a hydrocarbon radical, and they are all referable to the water type.

$H—O—H$,	$CH_3—O—H$,	$C_2H_5—O—H$,	$C_4H_9—O—H$.
Water.	Methyl Alcohol.	Ethyl Alcohol.	Butyl Alcohol.

If we replace two atoms of hydrogen which are not united to the same carbon atoms, by the radical HO, we form what is known as a **diatomic alcohol,** or **glycol.** Thus,

$$CH_3—CH_3 \text{ gives } HO—CH_2—CH_2—OH.$$

If three atoms of hydrogen united to three different carbon atoms be replaced by HO, a **triatomic alcohol,** or **glycerine** is produced. Thus—



$$CH_3—CH_2—CH_3 \text{ gives } HO—CH_2—\underset{\underset{\displaystyle H}{|}}{\underset{\displaystyle O}{\underset{|}{CH}}}—CH_2—O—H.$$

Tetratomic, pentatomic and hexatomic alcohols are known.

541. **Ethers.**—If, in a molecule of water, both the atoms of hydrogen be replaced by hydrocarbon radicals, an ether is produced. When the two radicals are alike, the compound is a **simple ether**; when unlike, it is said to be a **mixed ether.**

Examples of simple ethers—

$$CH_3—O—CH_3.$$

Methylic ether or methylic oxide.

$$C_2H_5—O—C_2H_5.$$

Ethylic ether or ethylic oxide.

Examples of mixed ether—

$$CH_3—O—C_2H_5.$$

Methyl-ethylic ether.

Haloid ethers, so-called, are the addition or substitution products containing one of the halogen elements; as,—

$$C_2H_5Br \text{ or } CH_3Cl.$$

Compound ethers are formed by replacing one of the hydrogen atoms in a molecule of water by a hydrocarbon radical, and the other by an acid radical; as,

$$C_2H_5—O—C_2H_3O.$$

542. **Aldehydes.**—This class of bodies are formed in the first stage of the oxidation of alcohols, and are intermediate between alcohols and acids. They differ from the corresponding alcohols by having two atoms of hydrogen removed; thus, ethyl alcohol, $C_2H_5—O—H$ forms $C_2H_3—O—H$, ethyl, or acetic aldehyde. Aldehydes by oxidation yield the corresponding acid. It will be observed that, as two atoms of hydrogen have been removed without putting any atoms in their places, aldehydes are unsaturated bodies. On this account, they are very liable to undergo changes with reagents. Hydrochloric and sulphuric acids and some other bodies convert acetic aldehyde, on standing, into two isomeric bodies called **paraldehyde** and **metaldehyde.** Ordinary aldehyde may be easily prepared by the slow or incomplete combustion of ether vapor. Heat a glass or iron rod and plunge it into a jar containing a little ether vapor, and the peculiar, pungent odor of aldehyde will be detected. Ordinary aldehyde is of little importance to the physician.

Paraldehyde, $C_6H_{12}O_3$, or $(C_2H_3OH)_3$, has been introduced lately as a substitute for morphia. It is a liquid at ordinary temperatures, but forms a white, crystalline solid at 10° C. (50° F.); it is but slightly soluble in water.

543. **Chloral** is an aldehyde in which the three hydrogen atoms of the radical are replaced by chlorine atoms; thus: $C_2H_3—O—H$ and $C_2Cl_3—O—H$.

Chloral is prepared by passing dry chlorine into absolute alcohol until saturated, and then adding sulphuric acid and distilling off the chloral. It is thus obtained as a colorless liquid, having a pungent irritating odor, and boiling at 94.5° C. (202° F.) It is a true aldehyde, which, by oxidation with nitric acid, gives **trichlor-acetic acid,** $C_2Cl_3O_2H$, a colorless, crystalline solid, which is soluble in water. This acid has been recommended as a test for albumen in urine. Chloral is converted into chloroform and a formate by the alkaline hydrates.

$$\underset{\text{Chloral.}}{C_2Cl_3OH} + \underset{\substack{\text{Sodium}\\\text{Hydrate.}}}{NaOH} = \underset{\text{Sodium Formate.}}{NaCHO_2} + \underset{\text{Chloroform.}}{CHCl_3}.$$

It unites with water to form **chloral hydrate,** $C_2Cl_3OH.H_2O$, a colorless, transparent, crystalline solid, having an aromatic, pungent odor and taste, and a neutral action on litmus paper. It is freely soluble in water, alcohol, ether, glycerine, benzine, fixed and volatile oils. When mixed with carbolic acid or camphor, it liquifies. It melts at about 58° C. (136.4° F.) and boils at about 95° C. (203° F.). It should be kept in glass stoppered bottles. Chloral hydrate is used in medicine under the name of chloral, to produce sleep. Its long continued use induces a chloral habit very difficult to cure.

Bromal, C_2Br_3OH, is also known.

544. **Ketones.** These bodies are the result of the first action of oxidizing agents upon **secondary alcohols**; *i. e.*, upon an alcohol which contains the group of atoms CHOH, instead of CH_2OH, as in ordinary alcohol.

By oxidation of these alcohols, the first result is to remove one atom of hydrogen, instead of two, as in the aldehydes.

Acetone, $(CH_3)—CO—(CH_3)$, is the best known ketone. The corresponding secondary alcohol would have the formula $CH_3—CHOH—CH_3$. Secondary alcohols, when oxidized, yield **ketones**, while primary alcohols yield **aldehydes.** Ketones, by further oxidation, break up and yield two acids, while aldehydes yield one acid with the same number of carbon atoms. Acetone sometimes appears in the urine of diabetic patients, and is detected by the addition of ferric chloride, with which it gives a deep brown-red color.

545. **Organic Acids.**—The characteristic feature of an organic acid molecule, is that it must contain the group —COOH (carboxyl), and the basicity of the acid will depend upon the number of these groups contained in its molecule. The organic acids partake of the general properties of the mineral acids. They may be referred to the water type and considered as one or more molecules of water, in which one atom of hydrogen has been replaced by an organic compound radical containing oxygen—a negative radical—while the other atom remains as replaceable, or basic hydrogen. As in inorganic acids, only those atoms of hydrogen which are linked to the carbon or negative group by oxygen are replaceable by a basic or metallic radical.

Acids may be formed by the oxidation of an alcohol, an aldehyde or a ketone. The presence of an alkali favors the formation of acids.

Acids may be **monobasic, dibasic, tribasic**, etc., according as their molecules contain one, two, three, etc., carboxyl groups (—COOH). Acids may also contain one or more hydroxyl groups, (OH), which are not basic, because not immediately attached to or associated with CO, to form the carboxyl group. The hydrogen of these hydroxyl groups is called **alcoholic hydrogen** to distinguish it from the other replaceable hydrogen, designated as **basic.** The number of hydroxyl groups in a molecule of either an alcohol or an acid is said to be its

atomicity. It is evident that the atomicity of an acid may be greater than its basicity, when it is said to be an alcohol acid.

Lactic acid, $(\mathrm{H{-}O{-}COCH(OH)CH_3} = \left(\mathrm{C_3H_4O} \left\{ \begin{matrix} \mathrm{OH} \\ \mathrm{OH} \end{matrix} \right.\right)$, is a good example of a diatomic and monobasic acid.

ALCOHOLS.

546. Having given a brief outline of the several groups of bodies which may be considered as derived from the hydrocarbons, viz., alcohols, ethers, aldehydes, ketones and acids, we shall now proceed to notice briefly those members which may be considered of most importance to the student of medicine; and we shall omit many of the compounds which are of chemical interest only.

Table of Monatomic Alcohols of the First Series, with corresponding Acids.

ALCOHOLS.	Formula.	Boiling Point.	FATTY ACIDS.	Formula.	Boiling Point.
Methyl Alcohol (Wood Spirit),	CH_3-O-H	66° C.	Formic Acid,	$HCHO_2$	100° C.
Ethyl " (Spirit of Wine),	C_2H_5-O-H	78.4° C.	Acetic Acid,	$HC_2H_3O_2$	118° C.
Propyl "	C_3H_7-O-H	97 ° C.	Propionic Acid,	$HC_3H_5O_2$	141° C.
Butyl "	C_4H_9-O-H	116 ° C.	Butyric Acid,	$HC_4H_7O_2$	163° C.
Amyl, or Pentyl Alcohol,.......	C_5H_{11}-O-H	137 ° C.	Valerianic, or Pentylic Acid,	$HC_5H_9O_2$	185° C.
Hexyl Alcohol, ...	C_6H_{13}-O-H	157 ° C.	Caproic, or Hexylic Acid,	$HC_6H_{11}O_2$	205° C.
Heptyl "	C_7H_{15}-O-H	176 ° C.	Œnanthic, or Heptylic Acid,	$HC_7H_{13}O_2$	224° C.
Octyl "	C_8H_{17}-O-H	195 ° C.	Caprylic, or Octylic Acid,	$HC_8H_{15}O_2$	233 ° C.
		Melting Point.			Melting Point.
Cetyl Alcohol, ...	$C_{16}H_{33}$-O-H	50° C.	Palmytic Acid,	$HC_{16}H_{31}O_2$	62° C.
			Stearic Acid,	$HC_{18}H_{35}O_2$	69.2°C.
Ceryl "	$C_{27}H_{55}$-O-H	79° C.	Cerotic Acid,	$HC_{27}H_{53}O_2$	78° C.
Melissyl "	$C_{30}H_{61}$-O-H	85° C.	Melyssic Acid,	$HC_{30}H_{59}O_2$	88° C.

METHYL ALCOHOL (Wood Spirit).

CH_3-O-H. Boils at 66° C. (150.8° F.)

547. **Methyl Alcohol** occurs among the products of the destructive distillation of wood, forming about one per cent. of the distillate. The alcohol is separated from the other products by careful fractional distillation. It is then redistilled after adding lime. This is then treated with dry calcium chloride, with which the alcohol combines to form a solid, crystalline body, from which other impurities may easily be separated; the alcohol is recovered from this solid by distilling with water. The water is finally removed by treating with quicklime, and distilling at

This reaction accounts for only 94 to 96 per cent. of the sugar employed, the remaining part being transformed into other products, such as succinic acid, glycerine, "fusel oil," etc. The yeast plant, under whose influence this change takes place, is known as **Torula, or Micoderma cerevisiæ.** The sugar to be fermented may be from the juices of fruit or sugar-cane, or it may be previously prepared from starch of the cereals by the aid of malt or sulphuric acid. The secondary, or side products in the fermentation vary somewhat with sugar from these various sources, and thus give rise to the different flavors of the fermented liquors. The alcohol may be separated from the water and other products after the fermentation by careful fractional distillation. If it be desired to prepare an alcohol containing more than 90 per cent., some substance must be added which will combine with and hold back the water, as quicklime, anhydrous copper sulphate, potassium carbonate, or barium hydrate, after which it is again distilled.

Absolute alcohol is a colorless, limpid liquid, of an agreeable odor and burning taste. It attracts water and mixes slowly with this liquid in all proportions, with the production of heat and contraction in volume. It boils at 78.4° C. (173° F.) and has never been solidified. It is neutral to test paper and burns readily with a non-luminous flame. It dissolves resins, essential oils, alkaline hydrates, calcium chloride and a large number of organic bodies. Oxidizing agents convert it into aldehyde and, finally, into acetic acid. Nitric acid (fuming) decomposes it very rapidly, giving a number of acids and ethers. If 90 per cent. alcohol is added to a solution of mercury or silver in nitric acid, a rapid ebullition takes place, with a crystalline deposit of **fulminate of silver** or **mercury**, which explodes by percussion, and is used in filling percussion caps. The formulæ are $CN.C(NH_2)Ag_2$ and $CN.C(NH_2)Hg$. The alkaline metals attack alcohol and give ethylates of the metals.

$$2C_2H_5OH + K_2 = 2C_2H_5OK + H_2.$$

Acids form with it compound ethers.

Physiological Action.—Alcohol is, when concentrated, a poison. Even when taken in large doses well diluted, it has frequently caused death, probably by paralysis of the muscles of respiration or of the heart. In full doses, it causes a feeling of warmth in the stomach, followed by congestion of this organ. After absorption, there is at first a feeling of exhilaration with exalted animal functions, quickened pulse and increased circulation, with dilatation of superficial blood vessels. This is often accompanied and always followed by incoherence of ideas and muscular actions, and finally, a general weakness of all the voluntary actions. The temperature is lowered, both in health and in most diseased conditions. The prolonged use of alcoholic beverages is characterized by general degenerative changes, either fatty or fibroid. There are no conclusive arguments to prove the benefit to be derived from the use of alcohol in health, but in certain diseased conditions it is of undoubted value.

Commercial Forms.—Commercial alcohol varies in strength from 85 to 95 per cent. The U. S. P. strength is 91 per cent. by weight, and 94 per cent. by volume. **Alcohol Dilutum** contains 45.5 per cent. by weight, or 53 per cent. by volume.

Tinctures are solutions of medicinal substances in alcohol.

551. **Alcoholic Beverages** are usually of three classes; viz., distilled, fermented and malt liquors. To the first class belong **spirits,** distilled from fermented liquors; **brandy,** from wine; **whisky,** from a mash of corn or rye; **rum,** from molasses; **gin,** from corn spirit flavored with juniper berries.

To the second class belong the various wines. To the third class belongs beer, prepared from malted barley and hops.* Ale, porter and stout differ from beer only in the selection and proportion of the malt, hops and flavoring materials.

We introduce here some results of the analyses of various distilled liquors, as found in the American market, for the purpose of comparison as to the ordinary strength of alcohol (N. Y. State Board of Health Report, 1882).

		Brandy.	Whiskey.	Rum.
Specific Gravity,		0.9297 to 0.9615	0.9018 to 0.9646	0.9126 to 0.9684
		Per cent.	Per cent.	Per cent.
Alcohol	By Volume,	30.80 to 50.40	28.90 to 60.30	26.40 to 50.30
	By Weight,	25.39 to 42.96	23.75 to 52.58	21.66 to 42.87

It seems, therefore, that there is a great variation in the strength of these liquors as ordinarily sold; and while they are supposed to contain from 40 to 50 per cent. of alcohol by volume, in reality, they usually contain from 35 to 45 per cent. It is, therefore, a very uncertain way of prescribing alcohol, to prescribe any one of these beverages. A much more certain method is to prescribe alcohol of known strength, flavored with ethereal essences, and softened with glycerine or syrup.

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* Glucose is now very generally used to take the place of part of the malt.

Wines contain from 6 to 25 per cent. of alcohol.

Port Wine	16.62 to 23.2	per cent.
Sherry "	16.00 to 25.0	"
Bordeaux, Red	6.85 to 13.0	"
" White	11.00 to 18.0	"
Champagne	5.80 to 13.0	"

Beers and porters contain from 1 to 10 per cent., average about 4 to 5 per cent. by volume. The average of extractive matters (dextrin, cellulose, sugar, lupulin and hop resin) is 4 to 15 in ale, 4 to 9 in porter, and about 5 per cent. in beer. All alcoholic beverages are subject to gross adulterations. Artificial beverages are frequently sold in all markets.

552. **Amyl Alcohol—Fusel Oil,*** **Potato Spirit,** $C_5H_{11}OH$. This alcohol is formed in small quantities at the same time with ethyl alcohol, during the fermentation of barley, corn, and especially potato mash. It is prepared from the residue left in the still, after the common alcohol is distilled off, by fractional distillation. The product coming over at 132° C. (269° F.) is that collected.

It is a colorless, oily liquid, possessing a peculiar, irritating odor which excites coughing, and a burning taste. It is not miscible with water, but mixes in all proportions with alcohol and ether. It is a good solvent of certain alkaloids. Taken internally it acts as a poison, both in the form of vapor and when taken by the stomach, producing dizziness, headache and intoxication. Much of the unwholesomeness of imperfectly rectified spirituous liquors arises from their contamination with fusel oil.

The principal use of fusel oil is as a source of amyl ethers, which are used extensively to prepare artificial flavoring extracts. Thus, the acetate has the odor of pears, and is used by confectioners under the name of "**pear oil,**" while the valerianate is used to give the flavor of apples, and called "**apple oil.**" The special description of the remainder of the monatomic and the diatomic alcohols, or **glycols,** will be omitted, as they are comparatively unimportant.

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* Fusel oil, properly speaking, is a mixture of several alcohols of which amyl alcohol is one.

TRIATOMIC ALCOHOLS, OR GLYCERINES.

553. **Ordinary Glycerine,** $C_3H_5(OH)_3$, was discovered by Scheele in 1779, and was called by him **the sweet principle of oils.** It is prepared on a large scale from the neutral fats, as a side product in the manufacture of soap and candles. These fatty bodies are composed of fatty acids in combination with glyceryl, the radical of glycerine; *i. e.*, they are compound ethers of the fatty acids and glyceryl. On treating these ethers with alkalies a salt of the alkali and acid is formed, termed a soap; the glycerine is set free by the reaction and remains dissolved in the water present. This process of decomposing a compound ether into an alcohol and acid or salt of the acid is called **saponification.** Neutral fats can also be saponified by treating them with superheated steam, which is the process now usually employed, in candle factories, at least. The glycerine is freed from the water by evaporation, and finally, by distillation with the aid of superheated steam. Glycerine is also formed during alcoholic fermentation, and is found in wines, etc. It is a colorless, syrupy liquid, possessing a sweetish taste and no odor; density 1.28 at 15.5° C. (60° F.). The officinal **glycerinum** has a density of 1.25. It is soluble in all proportions in water and alcohol, but not in ether. It absorbs water readily (hygroscopic) until it has absorbed twice its own volume. Its range of solubility is large, as will be seen by the table in the appendix. When heated in air, it boils at 290° C. (554° F.), and distills with partial decomposition. At low temperatures, it forms, under certain circumstances, a crystalline mass. When heated in the air to a high temperature, it takes fire and burns, leaving no residue. Boracic acid imparts a green color to its flame, or to a flame directed upon a platinum wire moistened with it. (This is one of the most convenient tests for boracic acid.) Chemically, glycerine is, as above stated, a triatomic alcohol; *i. e.*, it contains three hydroxyl groups. When oxidized, it yields glyceric acid. It is capable of undergoing some kinds of fermentation, as butyric fermentation. It unites with the alkalies and alkaline earths soluble in water, and thus prevents the precipitation of some of the metals by reagents which ordinarily precipitate them. It prevents the precipitation of copper hydrate by sodium or potassium hydrate, and has been recommended for this purpose in the preparation of Fehling's test solution for glucose.

Glycerine is sometimes adulterated with glucose, cane sugar syrup and water. The first and second of these will usually be easily detected by adding a solution of sodium hydrate (caustic soda) and enough copper sulphate to impart a blue color, and boiling for a few minutes, when the red cuprous oxide will be precipitated if these be present. The presence of water may be detected by taking the specific gravity, which should not be below 1.25. Glycerine is the basis of the manufacture of nitro-glycerine used in various forms of blasting agents, such as "dualine," "dynamite," "giant powder," "rend rock," etc., which are usually composed of nitro-glycerine and some porous substance in powder form.

Sulphuric acid combines with glycerine to form **glycero-sulphuric acid.** Glacial, or metaphosphoric acid forms **glycero-phosphoric acid.**

$$C_3H_5(OH)_2—PO_2=O=H_2.$$

This acid is one of the decomposition products of lecithin and protagon, two complex bodies found in nerve substance, especially of the brain. The acid itself has been found in the brain, nerves, muscles, yolk of egg, bile and pus. The phosphorus present in nerve matters probably exists in the form of either lecithin or protagon, both of which contain glycero-phosphoric acid. The salts of glycerine and organic acids will be referred to later. (See Neutral Fats.)

554. **Nitro-glycerine, or Glyceryl Trinitrate,** $C_3H_5\left\{\begin{matrix}NO_3\\NO_3\\NO_3\end{matrix}\right.$. —When glycerine is allowed to flow in a slow stream into a mixture of strong nitric and sulphuric acids kept cold by a freezing mixture, and the mixture afterward thrown into a large quantity of cold water, there separates out a heavy, colorless, poisonous oil—nitro-glycerine. It crystallizes at —20° C. (4° F.); sp. gr. 1.6. When inflamed in air, it burns quietly and rapidly; but when ignited by percussion or quick heating, especially in a confined space, it explodes with terrific violence, and hence is much used in blasting.

555. **Tetratomic Alcohols.**—None of this class of alcohols are of sufficient importance to be discussed here.

HEXATOMIC ALCOHOLS.

Formula, $C_nH_{2n-4}(OH)_6$.

556. In the hexatomic alcohols the six hydroxyl radicals are united to six different carbon atoms, so that there must be six carbon atoms in the nucleus. This class of alcohols include—

Mannite, $C_6H_8(OH)_6$.
Dulcite, $C_6H_8(OH)_6$.

557. **Mannite.**—This is the sweet principle of manna, and is found widely distributed in the vegetable kingdom. It occurs in celery, fungi and sea weeds, in the sap of the larch, the exuded sap of the apple, cherry, lime, etc., and the exuded sap of **Fraxinus Ornus,** which in the dry state forms commercial manna. It may be prepared artificially by acting upon grape sugar with sodium amalgam (nascent hydrogen), or by the so-called mucous and butyric fermentations of sugar.

To obtain it from manna, dissolve in half its weight of boiling water, add some albumen, to clarify, and filter through cloth. On cooling, the manna separates out. It may also be obtained pure by extracting manna with hot alcohol, and crystallizing. It forms fine, silky needles when crystallized from alcohol, but large, transparent, rhombic prisms from the aqueous solution. Mannite is intensely sweet, sparingly soluble in cold, readily soluble in hot water and alcohol, and insoluble in ether. It can readily undergo lactic and butyric fermentations, but not alcoholic. It unites with many metallic oxides, and also forms a large number of compound ethers. Mild oxidation produces **mannitose,** $C_6H_7O(OH)_5$. By further oxidation, it forms mannitic acid, $C_5H_6(OH)_5CO.OH$, and saccharic acid, $C_4H_4(OH)_4\left\{\begin{matrix}CO.OH\\CO.OH\end{matrix}\right.$.

558. **Dulcite.** This isomeride of mannite is found in **dulcose,** or dulcite manna, a crystalline substance from Madagascar, and in several plants. It has been prepared from milk sugar by treatment with sodium amalgam. The properties of dulcite resemble those of mannite.

Sorbite is a third isomeride of manna, found in the mountain ash.

CARBOHYDRATES.

559. Closely allied to the alcohols, and by some classed as alcohols, is the class of bodies known as carbohydrates. They contain six, or a multiple of six carbon atoms, and twice as many atoms of hydrogen as oxygen.

They may be divided into three groups:—

(1) SACCHAROSES.	(2) GLUCOSES.	(3) AMYLOSES.
$C_{12}H_{22}O_{11}$.	$C_6H_{12}O_6$.	$(C_6H_{10}O_5)n$.
+Cane Sugar.	+Dextrose, (Grape Sugar).	+Starch.
+Milk Sugar.		+Glycogen.
+Melitose.	—Levulose.	+Dextrin.
+Melizitose.	+Galactose.	—Inulin.
+Mycose.	—Sorbin.	Gums.
Synanthrose.	Eucalin.	Cellulose.
	Inosite.	Tunicin.

All these compounds occur ready formed, either in plant or animal organisms. The three groups are connected by a close relationship; indeed, those of 1 and 3 are easily converted into the glucoses by ferments or by boiling with dilute acids, by which water is added to them. The saccharoses and amyloses may be regarded as anhydrides of the glucoses. The chemical constitution of only a few of these bodies is known, and the molecular weight, with certainty, of none.

They all have the properties of polyatomic alcohols, and a few behave also like the aldehydes. The members of groups 1 and 2 dissolve in water, have a sweet taste and are called **sugars.** Most of them are **optically active,** or possess the power of rotating the plane of polarized light to the right (dextrogyrate) or to the left hand (lævogyrate). The right-handed substances are marked +, and the left-handed —.

(1) SACCHAROSES.

$C_{12}H_{22}O_{11}$.

560. **Cane Sugar.**—Cane sugar occurs in the juices of many plants, in most sweet fruits, in the nectar of flowers and in honey. It is found in large quantity in the juice of the sugar cane (Saccharum Officinarum), in sorghum (Sorgho Saccharatum), in beet-root, sugar-maple and in several species of palm.

Sugar is obtained principally from sugar cane, beet-root, sorghum and the maple, while some attempts have been made to obtain it from the common maize.

The juice of the sugar cane is expressed by passing the stalks between steel rollers; this is then mixed with milk of lime and heated to boiling to precipitate the coagulable materials, phosphates, etc.; these are removed, and the clear liquid evaporated, either in open pans or in vacuum pans under a reduced pressure, to a thick syrup, and then left to crystallize. The "raw sugar" is drained from the "mother-liquor," which is further evaporated and again left to crystallize. There is finally left a thick, brown liquid, called molasses, or treacle, containing uncrystallizable sugar.

The raw sugar is refined by dissolving in water, adding lime, and filtering through thick layers of bone charcoal to remove the color. It is then evaporated in vacuum pans and allowed to crystallize.

ne sugar crystallizes in large, transparent, monoclinic as are seen in rock candy. It is soluble in one-third its weight of cold, and in all proportions in hot water. It is slightly soluble in weak, but almost insoluble in absolute alcohol. It melts at 160° C. (320° F.), and solidifies, on cooling, to an amorphous mass, called "barley sugar."

The rotary power of sugar, in solution, is taken advantage of for its quantitative determination. The saccharometer is the instrument in common use for this purpose; its use depends upon the principle, that, in a solution of definite thickness, the rotary power of the solution is dependent upon the per cent. of sugar present, and by measuring the amount of rotation of the light, the amount of sugar in the solution can easily be determined.

When cane sugar is boiled with dilute sulphuric acid, it takes up water and is converted into a mixture of dextrose and levulose, called **inverted sugar.**

$$C_{12}H_{22}O_{11} + H_2O = CH_{12}O_6 + C_6H_{12}O_6.$$

The same change takes place under the influence of yeast, probably from some soluble ferment contained in it. When heated above its melting point, it gives off water, turns dark, and forms a bitter, brown, amorphous substance, called **caramel.** At a higher temperature it continues to lose water, decomposes, and gives off inflammable gases, consisting of marsh gas, and carbon monoxide and dioxide; at a higher temperature a distillate is obtained containing aldehyde, acetic acid, acetone, etc. Strong sulphuric acid chars sugar, leaving a black, voluminous mass. Dilute nitric acid oxidizes cane sugar to saccharic acid, then into tartaric, and, finally, oxalic acid. Concentrated nitric and sulphuric acids form with sugar a nitrate, $C_{12}H_{18}O_7(NO_3)_4$, which is amorphous and very explosive. On heating with sulphuric acid and manganic oxide, a large quantity of formic acid is formed. Cane sugar forms a series of metallic compounds called **sucrates,** or **saccharates.** An aqueous solution of sugar will dissolve calcium, barium, magnesium and lead oxides, and, with an alkali, copper and ferric oxides also. The barium salt crystallizes well.

MILK SUGAR.

$C_{12}H_{22}O_{11} + H_2O.$

561. This is an important ingredient of the milk of mammals, and is prepared principally from cows' milk by evaporating the whey after removing the cheese. Cows' milk contains from 4 to 5 per cent. of sugar, while woman's milk contains from 6 to 7 per cent. It crystallizes in large, hard prisms, has a feebly sweet taste, and is soluble in 6 parts of cold water. Yeast ferments

it with difficulty. Lactic and butyric fermentations take place readily. Oxidizing agents yield mucic and saccharic acids. It forms compounds, like cane sugar, with metallic oxides, and reduces alkaline copper solutions.

The remaining members of this group are unimportant.

(2) GLUCOSES.

$C_6H_{12}O_6$.

562. **Dextrose, or Grape Sugar.**—This sugar is widely diffused through the vegetable kingdom, occurring in grapes, honey, in most sweet fruits, sprouting grains (Malt) etc., usually mixed with an equal weight of levulose. It occurs in small quantities in the blood, yolk of eggs, and urine, and in larger quantity, in diabetic urine. It is manufactured on a large scale from corn starch by boiling with dilute sulphuric acid, neutralizing with lime, drawing off the clear syrup after settling, and evaporating it down to a thick syrup and allowing it to crystallize. Dextrose is less sweet than cane sugar, 2½ parts of the former giving the sweetening power of 1 part of the latter. It is frequently used to adulterate the light-brown varieties of cane sugar. When present in considerable quantities (5 per cent.) it may often be detected by its property of mashing between the teeth instead of crushing like cane sugar. Dextrose crystallizes with some difficulty, and, as usually met with in the market, it does not present to the naked eye a distinct crystalline appearance. It easily undergoes oxidation, especially in alkaline solutions, and thus acts as a reducing agent. It readily reduces silver, copper, bismuth and mercury salts in alkaline solutions. Silver, if ammonia be present, is reduced to the metallic state and deposits as a brilliant mirror on the surface of the vessel in which it is heated.

The ordinary methods of detection and estimation of dextrose depend upon its reducing power. Fehling's solution, which is in common use as both a qualitative and quantitative test for dextrose, is a solution of 34.64 grms. of pure crystallized copper sulphate, 173 grms. of Rochelle salt, and 80 grms. of sodium hydrate in a litre of distilled water. 1 c.c. of this solution is exactly reduced and decolorized by 5 milligrams of dextrose. For qualitative detection, alkaline solutions of copper, bismuth, indigo-carmine, picric acid or silver may be used. Dextrose undergoes alcoholic fermentation with great ease and is used in large quantities in making beer.

563. **Levulose.**—This sugar occurs mixed with dextrose in the natural sources of that sugar, mentioned above. It is obtained

with dextrose in inverted sugar, prepared by heating cane sugar with dilute mineral acids. It differs from dextrose in its rotary power, which is left-handed instead of right-handed; the other differences are of minor importance, except sweetening power, which is greater in this sugar than in grape sugar.

564. **Galactose** is obtained from milk sugar by inverting it with dilute sulphuric or hydrochloric acid. It is very fermentable, and reduces alkaline copper solutions.

Sorbin occurs in the berries of the mountain ash, and is not fermentable.

Inosite has been found in the muscular substance of the heart, in the lungs, liver, kidneys, etc., and in some plants. It is very sweet, and crystallizes in large, rhombic prisms, soluble in water. It does not ferment with yeast, reduce the alkaline copper solutions, or invert with acids or alkalies; it is optically inactive.

(3) AMYLOSES.

$(C_{12}H_{20}O_{10})n$.

565. **Starch, or Amylum.**—This body is found in nearly all plants. It is most abundant in the cereals, rice, potatoes and the seeds of plants. Starch occurs in the form of microscopical granules inclosed in the cells of the plant where they occur, very much as fat occurs in adipose tissue. The granules, examined under the microscope, are seen to be possessed of a distinct organized structure, which is different in each different kind of starch. They show, under the microscope, several concentric markings, arranged around a nucleus, or hilum, which is situated nearer one edge. The sizes and markings of the starches from the various sources vary sufficiently to admit of identification, an important fact in the detection of adulterations in foods.

Starch is prepared by grinding the grain or vegetable, and then suspending it in water, or spreading it on a sieve and running water upon it. By this means the starch granules are washed out of the cells and remain suspended in the water. This milky-looking liquid is allowed to settle, when the starch falls to the bottom; this sediment is taken out, dried, and sent into the market.

Wheat starch is often prepared by suspending the flour in water and allowing it to stand until the gluten is dissolved by putrefactive fermentation, when the starch may be washed and dried.

Starch is insoluble in cold water, alcohol and ether. Heated with water to a little above 60° C. (140° F.), the contents of the granules swell up, burst the envelopes and escape into the water. It appears at first as a very fine powder, but afterward disappears and forms an apparent solution, which, if concentrated, forms, on cooling, a jelly-like mass, called **starch paste.** On long boiling, starch enters into a soluble form. The same change is produced by a dry heat of 100° C. (212° F.), by diastase, or dilute sulphuric acid. With diastase it soon passes into the form of maltose and dextrin, and finally into dextrose. It forms metallic compounds with lead, lime and barium oxides. Soluble starch is precipitated by alcohol and solutions of subacetate of lead.

Starch dissolves in cold concentrated nitric acid; on the addition of water, **xyloidin,** a white powder, separates. The most characteristic reaction of starch is the dark blue color it forms with free iodine. This blue iodide of starch is easily decomposed, and disassociates when the solution is heated, but re-forms on cooling.

566. **Dextrin, or British Gum,** is an amorphous, yellowish-white, gum-like body, readily soluble in water. It is formed by heating starch above 150° C. (302° F.), or by the first action of malt diastase, or hot dilute sulphuric acid upon starch. This change is a progressive one. When diastase or pancreatic ferment is added to gelatinous starch (starch paste), it is liquefied in a few moments and becomes soluble starch, which is represented by the formula, $10(C_{12}H_{20}O_{10})$. This is gradually hydrated under the action of the ferment, giving rise to eight different steps in the process, and eight different dextrins:—

Soluble Starch, $10(C_{12}H_{20}O_{10}) + H_2O = 9(C_{12}H_{20}O_{10}) + C_{12}H_{22}O_{11}$ Maltose,
Erythro-dextrin, A, $9(C_{12}H_{20}O_{10}) + H_2O = 8(C_{12}H_{20}O_{10}) + C_{12}H_{22}O_{11}$ "
Erythro-dextrin, B, $8(C_{12}H_{20}O_{10}) + H_2O = 7(C_{12}H_{20}O_{10}) + C_{12}H_{22}O_{11}$ "
Acroo-dextrin, C, $7(C_{12}H_{20}O_{10}) + H_2O = 6(C_{12}H_{20}O_{10}) + C_{12}H_{22}O_{11}$ "

and so on until the final result obtained is—

Acroo-dextrin, F, $2(C_{12}H_{20}O_{10}) + H_2O = 2(C_{12}H_{20}O_{10})$ + no Maltose;

that is, the final result always leaves a portion of dextrin unchanged into maltose. The final result is about 20 per cent. of acroo-dextrin to 80 per cent. of maltose. Pancreatic diastase, however, has the power of slowly changing maltose to dextrose, and the lowest acroo-dextrin to maltose. Iodine gives a reddish color with dextrin.

567. **Glycogen.**—This substance resembles starch, and occurs in the liver of man and several animals, in the embryo, yolk of egg and some mollusks. It dissolves in cold water to an opalescent solution, but is insoluble in alcohol and ether. In many of

its properties it resembles dextrin. Its solutions are dextrogyrate. A diet of starches, sugars, inulin, glycerin, albumin of egg, fibrin and casein causes an increase of glycogen in the liver; while inosite, quercite, mannite, gums and fats fail to do so. Violent exercise, prolonged starvation, or fevers exhaust the liver of sugar. Glycogen of the liver is readily changed to acroodextrin, maltose and dextrose by a peculiar diastatic ferment found in the liver itself. Iodine gives a reddish-brown color with both dextrins and glycogen; but the color returns with the latter after discharging it with heat.

568. **Cellulose, or Lignin.**—This forms the principal part of the solid framework of plants. The pure substance may be prepared by treating raw cotton or linen fibre with potassium hydrate, acids, and ether, to remove foreign matters. It is a white solid, exhibiting the structure of the fibre from which it is obtained. It is soluble in a solution of cupric hydrate in ammonium hydrate. On the addition of an acid to this solution, it is precipitated as a white, amorphous mass, which, mixed with camphor and compressed into various articles, is sold under the name of celluloid. Cellulose is insoluble in water, alcohol, ether and all ordinary solvents. Strong sulphuric acid dissolves it, and on diluting with water, white flakes separate. When boiled with dilute sulphuric acid, it is converted into dextrine and dextrose.

When cotton-wool is steeped in a cold mixture of 1 part of strong nitric and 3 parts sulphuric acid for a few minutes, squeezed as dry as possible, placed in fresh acid for 48 hours, then pressed dry, and finally washed thoroughly in water, then in a weak solution of sodium carbonate, it possesses, when dry, great explosive properties, and is called **gun cotton.** The appearance and physical properties of the cotton remain unchanged, but it becomes a nitro-cellulose, the composition varying with the mode of preparation.

Pyroxylin is used as a source of **collodion**; it is a nitrocellulose, containing less NO_3 groups than gun cotton. It is prepared by dipping cotton into a mixture of 10 parts HNO_3 and 12 parts of H_2SO_4, allowing it to steep for 10 hours or until it is soluble in a mixture of alcohol and ether (1 to 3), then removing and washing it in cold water.

Collodion is prepared by dissolving 4 parts of pyroxylin in a mixture of 70 parts ether and 26 parts alcohol. **Flexible collodion** is a mixture of 92 parts of the above, 5 parts turpentine, and 3 of castor oil. **Styptic collodion** owes this property to 20 per cent. of tannic acid.

If dry unsized paper be dipped into a cold mixture of 2 parts of sulphuric acid and 1 of water for a few seconds, and then washed quickly in cold water containing a little ammonia, it is rendered very tough and strong, and is called **parchment paper.**

GUMS.

569. Although these bodies cannot be regarded as belonging to the polyatomic alcohols, they seem to be closely associated with them. The gums are amorphous bodies, more or less soluble in water, but insoluble in alcohol, and are converted into one of the glucoses by dilute sulphuric acid.

570. **Gum Arabic and gum-senegal** are derived from different kinds of acacia. They occur in rounded, irregular masses, which dissolve in water to form a thick, viscid solution.

These gums are composed of the compounds of potassium and calcium with **arabin.** Arabin treated with dilute H_2SO_4 yields arabinose, $C_6H_{12}O_6$, a non-fermentable glucose.

Cerasin is the insoluble part of the plum and cherry tree gums, which also contain arabin. **Gum-tragacanth** is the hardened juice of some kinds of *astragalus.* **Gum bassorin** is obtained from a species of cactus; it is insoluble in water, but swells up therein to a jelly-like mass. Many other plants furnish mucilages or gum-like bodies.

ETHERS.

571. **Ethers** are of three kinds; **simple, mixed** and **compound,** which have already been defined. (See Art. 541.)

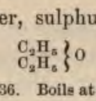
The simple ethers are usually prepared by the action of acids upon the corresponding alcohols. They bear the same relation to alcohols that oxides of the metals do to their hydrates.

Potassium hydrate, KOH. Potassium oxide, K_2O. Ethyl alcohol, C_2H_5OH. Ethyl ether $(C_2H_5)_2O$.

The simple ethers are mostly liquids, slightly soluble in water, soluble in alcohol, possess a peculiar odor, and when acted upon by alkalies, regenerate the alcohols.

ETHYL ETHER or ETHYL OXIDE.

SYNONYMS: Vinic ether, sulphuric ether, common ether.



$$\left.\begin{matrix}C_2H_5\\C_2H_5\end{matrix}\right\}O$$

Sp. gr. 0.736. Boils at 35.5°C. (96°F.)

572. **Common Ether** is prepared by heating a mixture of ethyl alcohol and sulphuric acid, and distilling over the resulting ether.

About 6 or 7 parts of commercial alcohol and 1 part of strong sulphuric acid are introduced into the retort, which is provided with two openings. Into one of these a thermometer is placed, while into the other is inserted a funnel tube. The retort is heated until the thermometer marks 130°C. (266°F.). Alcohol is now allowed to run in slowly through the funnel tube, while the temperature is kept between 130° and 140°C. (266° to 284°F.). The ether distils off with the water produced, and a small quantity of alcohol and sulphurous oxide. The crude ether floats upon the water as a distinct layer. To obtain it in a pure state, it is washed with dilute soda solution, dried over quicklime and calcium chloride, and redistilled by the heat of a water bath. By the above process, a small quantity of sulphuric acid may be made to etherize a very large quantity of alcohol.

The action takes place in two stages, as follows:

$$\underset{\text{Alcohol.}}{C_2H_5OH} + H_2SO_4 = \underset{\text{Ethylsulphuric acid.}}{(C_2H_5)HSO_4} + H_2O,$$

$$C_2H_5HSO_4 + C_2H_5OH = \underset{\text{Ether.}}{C_2H_5\text{—}O\text{—}C_2H_5} + H_2SO_4.$$

Properties.—Pure ether is a very volatile, mobile liquid, possessing a characteristic odor and burning taste. It is soluble in ten volumes of water and in all proportions in alcohol. It is highly inflammable, burning with a luminous flame; in handling it, therefore, care should be taken not to come near a flame. It dissolves resins, oils and many other organic bodies. It dissolves iodine, bromine, corrosive sublimate, sulphur and phosphorus. For anæsthetic purposes, it should not affect blue litmus; it should leave no residue when a quantity is evaporated on a watch glass, nor should it leave a foreign odor; it should boil in a test tube when the latter is held in the hand, and it should not impart a blue color to ignited copper sulphate (absence of water). A small quantity of alcohol, less than 6 per cent., is not a serious objection.

COMPOUND ETHERS.

573. There are a large number of these compounds known, but a few only are met with in medicine.

Acetic Ether ($C_2H_5—O—C_2H_3O$) is prepared by distilling a mixture of strong sulphuric acid, alcohol and sodium acetate. The distillate is washed with a solution of calcium chloride and milk of lime, decanted, dried over calcium chloride, and finally redistilled. Acetic ether is a colorless, limpid liquid, boiling at 74°C. (165°F.) and possessing a pleasant, fruity odor. It dissolves in about 10 parts of water, the water becoming acid from decomposition of the ether into acetic acid and alcohol. The refreshing smell of hock-vinegar and some old wines, is due to the presence of acetic ether. The ether is inflammable, burning with a bluish-yellow flame and acetous odor.

574. **Amyl Acetate** ($C_5H_{11}—O—C_2H_3O$) is prepared by distilling a mixture of amyl alcohol, sulphuric acid and sodium acetate. It has a pleasant, ethereal odor.

Ethyl benzoate, valerianate, butyrate, nitrite, nitrate, etc., are prepared very much in the same way as the above, substituting the salts of the acids here indicated for the sodium acetate.

575. **Ethyl Nitrite, or Nitrous Ether** ($C_2H_5—O—NO$) is a mobile liquid, boiling at 16.5°C. (61°F.), It has a sp. gr. of 0.947, and is insoluble in water, but freely soluble in alcohol. It is prepared by distilling a mixture of alcohol, potassium nitrite and sulphuric acid, or by gradually heating a mixture of equal parts of alcohol and strong nitric acid until it begins to boil; then remove the heat and allow it to distill slowly. It is purified as described above for the other ethers.

Sweet Spirit of Nitre is a mixture of ethyl nitrite, ethyl alcohol, aldehyde and acetic ether. It is prepared by adding together sulphuric acid, alcohol and nitric acid, and heating gradually to 80°C. (176°F.), when it is kept near this temperature while the ether distills over. This ether is now added to 19 times its weight of alcohol.

576. **Amyl Nitrite** ($C_5H_{11}—O—NO$) is prepared by passing nitrogen trioxide, N_2O_3, into amyl alcohol. It is a colorless or slightly yellow liquid, boiling at 99°C. (210.2°F.), and possesses the choking smell of amyl compounds generally. It boils at 96° C. (205° F.), specific gravity .872 to .874. Its vapor when inhaled, produces at first a sense of fullness and dizziness in the head; then flushing of the face, increased heart action, and lowering of the blood pressure and temperature. It may con-

tain, as impurities, nitric acid, amyl nitrate, amyl valerate, and hydrocyanic acid.

577. **Artificial Fruit Flavors.**—These flavoring extracts have come into use extensively in recent years, and are manufactured largely from various mixtures of compound ethers, organic acids and glycerin. The following formulæ will give some idea of the character of these mixtures: **Pine Apple** consists of chloroform, one part; aldehyde, one part; ethyl butyrate, five parts; amyl butyrate, ten parts, and glycerin, three parts. **Strawberry,** of ethyl nitrate, one part; ethyl acetate, five parts; ethyl formate, one part; ethyl butyrate, five parts; methyl salicylate, one part; amyl acetate, three parts; amyl butyrate, two parts; glycerin, two parts. **Pear,** of ethyl acetate, five parts; amyl acetate, ten parts; benzoic acid, one part, and glycerin, ten parts.

The ethers are to be dissolved in pure alcohol (specific gravity .83), and the numbers given indicate the quantity to be added to 100 parts of alcohol by measure. These mixtures, when taken in large quantities, produce deleterious effects; but, as the quantities actually used are very small, they probably have no such action. Besides the above mentioned and many other "fruit essences," these ethers are also extensively employed to improve the flavor of poor wines, and to fraudulently imitate wines, brandy, rum, whiskey, etc.

ORGANIC ACIDS.

578. An organic acid is built upon the water-type, and may be regarded as one or more molecules of water in which one-half the hydrogen is replaced by a compound organic radical containing oxygen; as $H—O—H$, $H—O—C_2H_3O$, $\left.\begin{matrix}H—O—\\H—O—\end{matrix}\right\}C_4H_4O_4$ (compare Art. 545). The number of organic acids known is very large; only a few of the most prominent ones can, therefore, be mentioned here.

Of the many series of acids, the most numerous and important are those of the **acetic** or **fatty acid** series, corresponding to the marsh gas series of hydrocarbons.

The acids of this series are obtained by the oxidation of the corresponding alcohols or aldehydes.

$$\underset{\text{Alcohol.}}{C_2H_5—O—H} + O_2 = \underset{\text{Acetic Acid.}}{C_2H_3O—O—H} + H_2O.$$

$$\underset{\text{Aldehyde.}}{2C_2H_3—O—H} + O_2 = \underset{\text{Acetic Acid.}}{2C_2H_3O—O—H}.$$

They may also be obtained by the action of the alkaline hydrates upon the cyanides of the radical of the next lower alcohol.

$$CH_3Cy + KOH + H_2O = NH_3 + KC_2H_3O_2.$$

A few have been obtained by synthesis from carbon monoxide.

$$CO + KOH = \underset{\text{Formic Acid.}}{CHO—O—K.}$$

THE FATTY ACID SERIES.

MONATOMIC ACIDS.

$$C_n H_{2n}O_2 = C_n H_{2n-1}O—O—H.$$

Of the large number of acids in this group, we shall notice formic, acetic, butyric, valeric, palmitic, margaric and stearic.

FORMIC ACID.

$CHO—O—H.$

579. **Formic Acid** is a colorless liquid, of a very acid reaction and sharp, pungent odor. It boils at about 100° C. (212° F.) and solidifies at about 0° C. (32° F.). It exists ready formed in the red ant, stinging nettle and pine needles. It acts as a reducing agent, reducing silver and mercury salts and depositing the metals. It is used in silvering glass to reduce the silver, which deposits upon the walls of the containing vessel. The best method of obtaining it is by heating to about 100° to 110° C. (212° to 230° F.) pure anhydrous glycerin and dry oxalic acid, adding more oxalic acid from time to time, and continuing the distillation. Carbon monoxide dissolved in potassium hydrate, yields some potassium formate.

ACETIC ACID.

$C_2H_3O—O—H.$

580. **Acetic Acid** occurs frequently in the form of acetates in some vegetable and animal fluids. It is usually obtained by the fermentation of saccharine fluids after they have undergone the alcoholic fermentation, or by the dry distillation of wood, starch, etc.

In distilling wood, gases, methyl alcohol (wood spirit), acetic acid, water, creasote and tar are obtained. The liquid portion is distilled at a gentle heat, when the alcohol is separated. The remaining liquid, containing the acetic acid, is saturated with sodium carbonate, evaporated to dryness, and heated to 250° or

350° C. (482° to 662° F.) to char the tarry matters. The residue containing sodium acetate is dissolved in water, filtered, and allowed to crystallize out. If the free acid is desired, the residue, after carbonizing, is distilled with a slight excess of sulphuric acid. This gives a colorless, strongly acid, sour-smelling liquid, which crystallizes at about 17° C. (63° F.) and is known as glacial acetic acid.

Acetic acid applied to the skin blisters, and causes considerable pain. It acts as a styptic when not too strong. It is soluble in water, alcohol and ether in all proportions. It dissolves resins, camphor, fibrin and coagulated albumin. It precipitates mucin, and is used to separate this body from its solutions. It is also used, with the aid of heat, as a test for albumin; but care is taken not to add too much, as it dissolves the albumin. Under the action of chlorine, acetic acid furnishes a series of chlorine substitution compounds, in which the chlorine is substituted for the hydrogen of the radical; thus, we have monochlor, dichlor and trichlor-acetic acids.

$$C_2H_2ClO_2H, C_2HCl_2O_2H, \text{ and } C_2Cl_3O_2H.$$

The last of these is mentioned elsewhere as a test for albumin in urine.

581. **Vinegar.**—This name is given to the mixture obtained by the fermentation of wine, cider, whiskey, molasses, infusion of malt, etc., under the influence of the growth of **mycoderma aceti**, and should contain at least four per cent. of acetic acid. Alcoholic fermentation always precedes the acetous.

As vinegar always contains more or less of this ferment, called mother of vinegar, it is customary to add some of this fluid to start the process. The fermentation takes place slowly, in vats or casks, because of the small amount of surface presented to the air. The process is rendered very much more rapid by allowing the fluid to trickle over beech-wood shavings or chips placed upon trays or in perforated barrels, so as to expose a large surface to the air. After having passed over the shavings four times, it will be found to be pretty thoroughly acetified. The temperature should be kept at about 25° C. (77° F.).

The vinegar of the market is frequently adulterated with 1, water; 2, mineral acids, especially sulphuric; 3, metallic impurities, as arsenic, lead, zinc, copper and tin; 4, wood vinegar; 5, organic substances, such as coloring matters, capsicum, etc. The addition of water can only be detected by the estimation of the per cent. of acetic acid.

The most objectionable adulterant is sulphuric acid. The simplest method of detecting free mineral acids is to evaporate a portion of the vinegar to dryness; then heat the residue to dull redness for some time, to convert the acetates of the alkaline metals into carbonates, which can easily be detected by their effervescence with hydrochloric acid. If any free mineral acid existed in the vinegar, it would expel the acetic acid from the alkalies and convert them into sulphates, which remain unchanged on ignition.

Another test for mineral acids is methyl violet. Two or three drops of a solution of this compound (0.1 to 100) added to 25 c. c. of vinegar, remains unchanged if pure. If .2 per cent. of any mineral acid be present, the color is blue; if .5 per cent., blue-green; if 1 per cent., green.

A simple test for sulphuric acid is to evaporate a portion to dryness in a white porcelain dish, with a little cane sugar. At the end of the process, the residue becomes black by the charring of the sugar by the acid. A small quantity of sulphuric acid is sometimes added to make the vinegar keep. The poisonous metals likely to be found are mercury (corrosive sublimate), copper, arsenic and lead. These metals may be detected by saturating the vinegar with hydric sulphide, or by separate tests for each. Burnt sugar, capsicum, etc., may be detected by taste or odor in the residue left on evaporation. The acetates are eliminated from the body as carbonates.

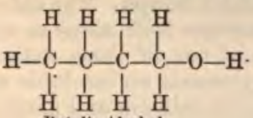
BUTYRIC ACID.

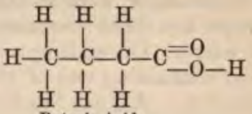
$C_4H_7O.O—H.$

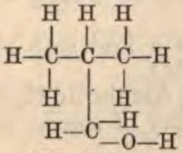
582. This acid is found with other fatty acids in butter, human perspiration, fæces, in flesh juice, and in some beetles. It exists in butter as a glyceride or glyceric ether. Pathologically, it appears in a free state in urine, blood, and ovarian cysts, and in the sputa of gangrene of the lung and bronchiectasis. It also appears in the intestinal contents as the result of secondary fermentation of saccharine articles of food.

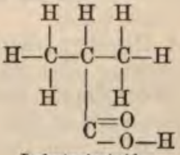
It is best prepared by maintaining at a temperature of 35 to 40°C (95° to 104°F.) a solution of sugar containing lime or chalk, and sour milk or rotten cheese. A mixture of 10 parts of sugar, 1 part of cheese, and 10 parts of chalk answers very well. Lactate of calcium is first produced, which afterward changes, under the influence of the ferment of cheese, into butyrate of calcium. The solution should remain alkaline or neutral. Carbon dioxide and hydrogen are set free. When the fermentation is finished, 30 or 40 parts of crystallized sodium carbonate is added, and the mixture is warmed and filtered. The filtrate is evaporated nearly to dryness, and hydrochloric or sulphuric acid added, which sets free the butyric acid as an oily layer, which may be purified by distillation. It boils at 162°C. (323°F.). It is a colorless liquid with the characteristic penetrating odor of rancid butter. It is soluble in pure water, but separates if soluble salts are added to the solution. It is soluble in alcohol, oils and ether. The butyrates are all soluble in water.

Isobutyric acid, an isomer of the foregoing, is obtained by oxidation of secondary or isobutylic alcohol. The following graphic formulæ show the constitution of these bodies:—


Butylic Alcohol.


Butyric Acid.


Isobutyl Alcohol.


Isobutyric Acid.

VALERIC, or VALERIANIC ACID.

$C_5H_9O.OH$.

583. Occurs in valerian root, and is obtained by distilling the powdered root with water. It is best prepared by the oxidation of amylic alcohol. A mixture of one part of amyl alcohol and four parts of concentrated sulphuric acid is run slowly into a retort containing four parts of water and five of potassium dichromate. The first product is valeraldehyde, which distills over. By elevating the beak of the retort so as to run the aldehyde back into the oxidizing mixture, it is changed into valerianic acid. The mixture is finally distilled, the distillate neutralized with sodium carbonate, evaporated and decomposed with sulphuric acid.

Valerianic acid is a thin, oily liquid, boiling at 175°C. (347°F.), and possessing a sour, old cheese odor. The most of the valerianates are soluble in water, and, when moist, smell like the acid.

Valerianates of ammonium, bismuth, caffeine, quinine, iron and zinc are used in medicine.

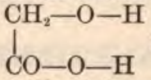
OLEIC ($C_{18}H_{34}O_2$), PALMITIC ($C_{18}H_{36}O_2$), AND STEARIC ($C_{16}H_{32}O_2$) ACIDS.

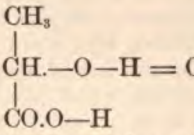
584. These three acids exist as glycerides in most fats and oils. Palmitic acid melts at 62°C. (143.6°F.), stearic at 69°C. (156°F.), and oleic is liquid at temperatures above 4°C. (39°F.). When once solidified, this last acid melts at 14°C. (57°F.). Neither of these acids distills without decomposition.

The acids are separated from the glycerin by **saponification**; *i. e.*, by treatment with caustic alkalies, or, still better, by superheated steam (see Glycerin). **Stearin candles** are composed of a mixture of stearic and palmitic acids. The salts of these three acids with metals, are called **soaps.** The only soluble soaps are those containing potassium (soft soap) or sodium (hard soap). On adding a solution of one of these soaps to a solution containing one of the heavy metals, an insoluble soap is precipitated. All are familiar with the calcium and magnesium soaps, which form as a curdy precipitate on adding ordinary soap to hard water.

Lead Plaster (diachylon plaster) is a lead soap, prepared by saponifying olive oil with litharge, PbO.

585. **Monobasic Acids containing Alcoholic Hydroxyl.**—But two acids of this group are worthy of special mention; viz., **glycollic** and **lactic acids.**

Glycollic Acid,  $\begin{matrix}\mathrm{CH_2—O—H}\\ |\\ \mathrm{CO—O—H}\end{matrix}$ is found in unripe grapes, and in the green leaves of Virginia creeper. When pure, it forms large, regular crystals, which deliquesce in moist air, and melt at 79° C. (174° F.).

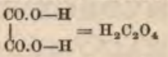
Lactic Acid,  $\begin{matrix}\mathrm{CH_3}\\ |\\ \mathrm{CH.—O—H}\\ |\\ \mathrm{CO.O—H}\end{matrix}$ $= C_3H_6O_3$ is the acid of sour milk and sour cabbage, and is produced by a special ferment, called **lactic ferment.** It is found in small quantity in the gastric juice, urine and intestinal juices.

It is produced on a large scale by lactic fermentation of cane sugar and glucose. Flour is first treated with dilute sulphuric acid, to convert the starch into glucose; the free acid is then neutralized with milk of lime. To this is then added sour milk, and it is allowed to ferment; care being taken to stop the process before butyric acid fermentation sets in, by heating the mixture to boiling. The hot solution of calcium lactate is separated by filtration, evaporated down, and allowed to crystallize. From this salt, the acid may be obtained by saturation with sulphuric acid. A solution containing 75 per cent. of the acid is officinal in the U. S. P. It is a colorless, syrupy, odorless, very acid liquid, freely miscible with water, alcohol and ether. It is used in the preparation of syrup of the lactophosphates, U. S. P.

Ferrous Lactate is used in medicine. It occurs in pale, greenish-white, odorless, crystalline crusts or grains, permanent in air, and soluble in water.

DIBASIC ACIDS.

OXALIC ACID.

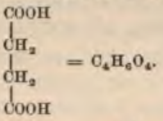


$$\begin{array}{l} CO.O-H \\ | \\ CO.O-H \end{array} = H_2C_2O_4$$

586. This important acid exists in many plants (Oxalis, Rumex, Rhei) as calcium or potassium oxalates. Calcium oxalate is also found in the urine.

Oxalic acid is easily obtained by acting upon many organic substances with oxidizing agents. Glycol, glycollic or acetic acids may be made to yield it. It is best prepared from the carbo-hydrates (sugar, starch, etc.) by treating them with strong nitric acid or by fusing with caustic potash. Commercial oxalic acid is prepared by fusing sawdust with a mixture of caustic soda and potash. The acid crystallizes in colorless prisms with two molecules of water, which they lose at 100° C. (212° F.). It is soluble in eight parts of cold water, and in alcohol. It is very sour and poisonous. Antidotes are chalk, lime, or magnesia. On heating the acid it is resolved into carbon monoxide, dioxide, water and formic acid. Strong oxidizing agents convert it into carbon dioxide and water. It is a strong dibasic acid, and forms both acid and neutral salts with most of the metals.

587. Succinic Acid—Acidum Succinicum.

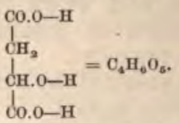


$$\begin{array}{l} COOH \\ | \\ CH_2 \\ | \\ CH_2 \\ | \\ COOH \end{array} = C_4H_6O_4.$$

This acid is found ready formed in amber and some other resins, in several plants, and in the animal organism. It is one of the products of alcoholic fermentation of sugar, and may be prepared by the action of reducing agents on malic and tartaric acids. It may be obtained in quantity by the dry distillation of amber; the aqueous portion of the distillate is heated to boiling, and filtered; on cooling, crude succinic acid crystallizes out. Succinic acid crystallizes in monoclinic prisms, melting at 180° C. (356° F.), and decomposing into water and succinic anhydride at 235° C. (455° F.) It is soluble in 23 parts of cold water, and very freely soluble in hot water. By adding neutral solution of ferric chloride to a soluble succinate, a brown, gelatinous, ferric succinate is produced; this is a reaction used as a qualitative test for the acid. The succinates of the alkaline metals are soluble; those of the other metals are either slightly soluble or insoluble.

DIBASIC ACIDS CONTAINING ALCOHOLIC HYDROXYL.

MALIC ACID.

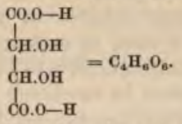


$$\begin{array}{l} CO.O-H \\ | \\ CH_2 \\ | \\ CH.O-H \\ | \\ CO.O-H \end{array} = C_4H_6O_5.$$

588. **Malic Acid** occurs in many acid fruits, as cherries, apples, raspberries, gooseberries, rhubarb (stalks and leaves), unripe mountain ash berries, unripe grapes and quinces. It is best prepared by nearly saturating the boiled and filtered juice of the berries of the mountain ash with milk of lime. On continued boiling, calcium malate, $CaC_4H_3(HO)O_4.H_2O$, separates as a crystalline powder, from which the acid may be obtained by treatment with dilute nitric acid. Malic acid crystallizes in groups of small, colorless, deliquescent crystals. It melts at 100° C. (212° F.), and decomposes at 150° C. (302° F.). Putrefactive ferments convert malic into acetic, succinic, butyric and carbonic acids.

There are three isomeric malic acids possible, of which two are known. The alkaline malates are soluble; other malates are slightly soluble or insoluble; all are crystalline.

TARTARIC ACID.



$$\begin{array}{l} \mathrm{CO.O{-}H} \\ | \\ \mathrm{CH.OH} \\ | \\ \mathrm{CH.OH} \\ | \\ \mathrm{CO.O{-}H} \end{array} = C_4H_6O_6.$$

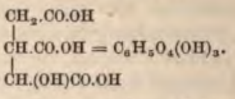
589. **Tartaric Acid** is a dibasic, tetratomic acid. (Compare the formulæ of succinic, malic, and tartaric acids.) Four isomeric, tartaric acids are known; two of which—dextro and lævo-tartaric acids—are optically active, and two—racemic and mesotartaric acids—are optically inactive. Ordinary, or dextro-tartaric acid is found in many fruits, particularly in ripe grapes, as acid potassic tartrate (**cream of tartar**), which, during the fermentation of the must, is deposited, mixed with yeast, coloring matter, calcium tartrate, etc., as a brown crust, or deposit, known as crude **argol.** Tartaric acid is prepared from argol by first treating with hot water, filtering, decolorizing with animal charcoal, converting the acid potassium tartrate into calcium tartrate by the addition of milk of lime, then decomposing this with sulphuric acid. Tartaric acid is thus obtained in solution, and may easily be separated by crystallization. Tartaric acid is usually made in the same factories where **cream of tartar** is prepared in large quantities. It usually occurs in beautiful, oblique prisms, permanent in the air, soluble in one-half their weight of water, or in 2.5 parts of alcohol, and insoluble in ether. When heated, it melts at 180° C. (356° F.), and forms metatartaric and pyrotartaric acids, and tartaric anhydride ($C_4H_4O_5$). At higher temperatures it decomposes with a burnt-sugar odor.

Tartaric acid has a strong, acid taste. It precipitates calcium in neutral or alkaline solution, but not in strong acid solutions. Ammonium salts prevent this precipitation. Heated with hydriodic acid and phosphorus, tartaric is first changed into malic, and then succinic acids.

The principal tartrates are the neutral and acid potassium tartrates, sodio-potassium tartrate (**Rochelle Salt**) and **tartar emetic,** or antimonyl potassium tartrate, all of which are mentioned in another place. Ferro-potassium tartrate and ferro-ammonium tartrate are also used in medicine. They both belong to the class of substances known as **scale compounds**; *i. e.*, compounds which do not crystallize readily and are prepared by spreading the material, evaporated to a syrup, upon plates of glass to dry, and then scraping off the thin scales. The above mentioned compounds occur in the form of garnet-red scales, slightly deliquescent, very soluble in water, but insoluble in alcohol. Tartaric acid is used in making effervescing drinks, in calico printing, and by confectioners, to prevent the crystallization of the sugar. When taken in too large quantities, it acts as an irritant poison. One ounce has caused death.

TRIBASIC ACIDS.

CITRIC ACID.



$$\begin{array}{l} CH_2.CO.OH \\ | \\ CH.CO.OH = C_6H_5O_4(OH)_3. \\ | \\ CH.(OH)CO.OH \end{array}$$

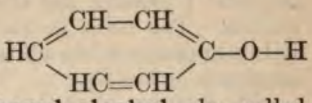
590. This acid occurs in the juice of lemons, currants, gooseberries, beet roots and other plants. It is manufactured on a large scale from lemon juice, which is clarified by boiling it with albumin, and then saturated, while hot, with powdered chalk or milk of lime. The precipitated calcium citrate is decomposed by an equivalent quantity of sulphuric acid, and filtered from the resulting gypsum. On evaporating the filtrate, the acid crystallizes out in large, transparent, rhombic prisms, having an agreeable, sour taste, and containing one molecule of water of crystallization. The acid melts at 100° C. (212° F.), and is readily soluble in water and alcohol. At 175° C. (347° F.), the acid loses water, and is converted into **aconitic acid,** $C_6H_3O_3(OH)_3$.

Solutions of citric acid soon develop mould, and are thereby decomposed. Citric acid forms three classes of well defined salts with the metals; citrates of the alkaline metals are soluble in water. The citrates are decomposed into carbonates in the body, and, in the case of the citrates of the alkalies, are eliminated by the kidneys as carbonates; hence, these citrates are frequently prescribed in acid conditions of the urine. Citric acid is not known to exert any injurious action upon the economy, even in considerable quantities.

Citrates of bismuth, iron, iron and ammonium, iron and quinine, iron and strychnine, lithium, potassium, bismuth and ammonium, and syrup of citric acid are officinal in the U. S. P. When boiled with excess of lime water, citric acid precipitates basic calcium citrate. This distinguishes it from oxalic and tartaric acids.

ACIDS OF THE BENZOL SERIES.

591. **Phenic, or Carbolic Acid,**



$= C_6H_5OH$, which is, in reality, **phenyl alcohol**, also called phenol, is met with in the products of the destructive distillation of coal and wood, and is found in small quantities in human urine, and in castoreum. In the arts, it is prepared from that portion of the coal tar oil, distilling between 150° C. (302° F.) and 190° C. (374° F.). This is agitated with a concentrated solution of caustic soda, when a crystalline carbolate of sodium is formed, while the neutral oils are left unacted upon. After the latter have been separated, the carbolate is decomposed by hydrochloric acid, and the impure carbolic acid thus obtained again treated with sodium hydrate. On exposing this solution to the air, the greater portion of the impurities become oxidized, and separate as a tarry mass. The clear solution of sodium phenate (or carbolate) is again decomposed by hydrochloric acid, and the resulting carbolic acid separated and submitted to distillation. From the portion passing over below 190° C. (374° F.) phenol separates out as colorless needles on cooling, which melt at 42° C. (108° F.) and boil at 184° C. (364° F.).

Pure phenol is a crystalline solid, having a characteristic odor and pungent, caustic taste, producing a white eschar with animal tissues. The crystals are liquefied by the addition of about 5 per cent. of water. This liquid is made turbid by the addition of more water, until 2000 parts are added, when the acid dissolves to a clear solution. It is soluble in twenty parts of water at the ordinary temperature. The addition of glycerin increases its solubility. It is readily soluble in alcohol, ether, benzol, carbon disulphide, glycerin, and fixed and volatile oils. Carbolic acid coagulates albumin, and its aqueous solution gives a permanent violet-blue color with ferric chloride, while that produced by creasote changes to green and then brown. It forms a white precipitate with bromine water. When quite pure, carbolic acid is permanent in the air; but the commercial acid frequently changes to a pink or red color. Ammonia and chlorinated soda solution produce a blue color with carbolic acid. Carbolic acid is very much used as an antiseptic in medicine and in the arts. Some of the carbolates have also been used for the same purpose. Carbolic acid is very poisonous when taken into the body, and cases of fatal poisoning by it are not uncommon.

Dangerous symptoms have been produced by six or seven drops, and fatal poisoning has occurred from its use as a dressing of wounds. The urine in such cases is dark-colored and smoky, and its appearance should be watched while using the acid, either internally or locally.

592. **Sulpho-carbolates.**—If one part of crystallized carbolic acid is mixed with one part, by weight, of strong sulphuric acid, phenyl-sulphuric or sulpho-carbolic acid is formed.

If this solution is diluted with water and barium carbonate added in excess, a solution of barium sulphocarbolate is produced. From this, other sulphocarbolates are obtained by double decomposition. The sulphocarbolates are all soluble in water, glycerin, and alcohol. The following salts have been used in medicine:—**Potassium Sulpho-carbolate,** $KC_6H_5SO_4.H_2O$, crystallizing in shining needles; **Calcium Sulpho-carbolate,** $Ca(C_6H_5SO_4)_2.6H_2O$, which forms scaly crystals; **Magnesium Sulpho-carbolate** and **Zinc Sulpho-carbolate,** crystallizing in rhombic prisms. These salts may be taken internally in doses of 10 to 30 grains. They prevent fermentation in organic solutions.

PICRIC, OR CARBAZOTIC ACID.

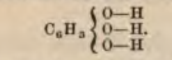
$C_6H_3(NO_2)_3—O—H.$

593. When phenol and a number of other bodies containing the benzol nucleus, such as salicin, indigo, balsam of Peru, silk, wool, etc., are treated with strong nitric acid, a yellow, very bitter, crystalline, trinitro-phenol is obtained, which is known as picric acid. It is used as a yellow dye for silk and wool, and as a test for albumen in urine. When quickly heated, it detonates. It is soluble in alcohol, and slightly so in water.

594. **Dihydro-phenol, or Hydroquinone,** $C_6H_4\left\{\begin{matrix}OH\\OH\end{matrix}\right.$, is of little importance to the physician.

Resorcin, $C_6H_4\left\{\begin{matrix}NO\\OH\end{matrix}\right.$, has come into prominence recently in the treatment of certain maladies. It forms colorless crystals, melting at 99°C. (210°F.), and boiling at 271°C. (519°F.). It is very soluble in water, and possesses a sweetish, harsh taste. It is prepared by fusing certain gum resins (galbanum), extract of sapin-wood, or Brazil-wood, with caustic potash. It is a strong anti-fermentative.

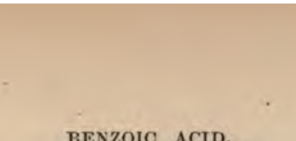
PYROGALLIC ACID.



$$C_6H_3\left\{\begin{matrix}O-H\\O-H.\\O-H\end{matrix}\right.$$

595. This acid is a triphenol, and is prepared by heating gallic acid, $C_6H_2(OH)_3CO_2H$, to 300°C. (572°F.). It crystallizes in white, shining needles, and gives a dark-blue color with ferrous, and red with ferric salts. An alkaline solution, when exposed to the air, rapidly absorbs oxygen and assumes a dark color. Silver and gold salts are reduced, in the light, to the metallic state by it; hence it is used as a developer in photography.

BENZOIC ACID.



596. This acid occurs in **gum Benzoin** and in the leaves of the aspen; also found in the urine of herbivorous animals. Benzoic acid is prepared from gum benzoin by sublimation. The powdered resin is gently heated in an iron pan covered with a perforated paper cover; over this is placed a paper or felt cone; the benzoic acid is sublimed and collects on the inner surface of the cone. It is also prepared by boiling the resin with milk of lime, filtering, concentrating the filtrate by evaporation, and precipitating the acid with hydrochloric acid. It is also prepared from naphthalene, $C_{10}H_8$, occurring in coal tar, and from hippuric acid found in the urine of cows.

Properties.—Benzoic acid occurs in large, thin, brilliant plates, or needles; it melts at 120°C. (248°F.) and boils at 250°C. (482°F.), but sublimes at a lower temperature. It is soluble in hot water and alcohol, but sparingly soluble in cold water. It has a peculiar, aromatic, irritating odor.

The **benzoates** are generally soluble in water.

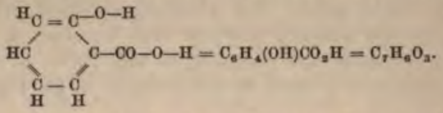
The benzoates of ammonium, lithium and sodium are officinal. The one most used is ammonium benzoate. This salt occurs in colorless, transparent, prismatic crystals, or in white and granular crystals, becoming yellow on long exposure to the air. It has a slight odor of the acid, a saline, afterward bitter taste; soluble in 5 parts of water.

The sodium salt is soluble in 1.8 parts of water, has a sweetly astringent taste, is free from bitterness, and has a neutral reaction.

The acid and its salts have an antifermentative and antiseptic action.

Solutions of benzoates give a flesh-colored or reddish precipitate with ferric chloride.

SALICYLIC, OR OXYBENZOIC ACID.



$= C_6H_4(OH)CO_2H = C_7H_6O_3$.

597. This acid occurs in the flowers of several species of *spirœa*, and its methyl ether, $C_6H_4\left\{\begin{matrix}OH\\COOCH_3,\end{matrix}\right.$ forms a large part of oil of wintergreen, from which it was formerly prepared. It is now prepared on a large scale by passing CO_2 into a heated retort containing sodium carbolate.

$$C_6H_5O.Na + CO_2 = C_6H_4(OH)CO.ONa,$$

When pure, salicylic acid occurs as fine, white, prismatic, needle-shaped crystals, permanent in air, having a sweetish, slightly acrid taste and acid reaction; soluble in 450 parts of cold water, and in 2.5 parts of alcohol. Its solutions give an intense violet color with ferric salts. This is used as a test.

The **salicylates** of lithium, sodium and some of the alkaloids are used in medicine. The **salicylic ethers** are very numerous.

598. Closely related to salicylic acid is a glucoside of oxybenzoic alcohol, called **salicin.**

$$C_6H_4\left\{\begin{matrix}OH\\CH_2OH.\end{matrix}\right. \qquad C_6H_4\left\{\begin{matrix}OH\\CH_2O(C_6H_{11}O_5).\end{matrix}\right.$$

Oxybenzoic Alcohol. Salicin.

Salicin is found in the bark of different kinds of willow and poplar, and in castoreum. It is prepared by boiling the bark with water and lead oxide, filtering, precipitating the lead with hydric sulphide, and evaporating down. The salicin separates from the solution in fine, colorless, needle-like prisms, having a very bitter taste and neutral reaction; soluble in 28 parts of water, and in 30 parts of alcohol; insoluble in ether or chloroform; very soluble in hot water. It is used in medicine.

HIPPURIC ACID.

$C_9H_9NO_3$.

599. This acid is found in the urine of herbivorous animals, and in small quantity in human urine. When benzoic acid or oil of bitter almonds is taken internally, hippuric acid appears in the urine. It is generally prepared by evaporating the urine of the cow or horse to one-tenth its volume, and precipitating with hydrochloric acid. It forms large, rhombic prisms, soluble in hot but sparingly in cold water.

URIC ACID.

$C_5H_4N_4O_3$.

600. Uric acid is found in the urine of all animals, and in the excrement of birds, as ammonium urate. Human urine in the healthy state contains but a small quantity of the acid, but it is increased in most febrile disorders, in acid dyspepsia, and in diseases of the lungs.

As uric acid is very sparingly soluble in water, it often crystallizes out as a sediment, and in some gouty subjects, the acid sodium salt crystallizes out in certain regions in the vicinity of the joints.

Uric acid is a white, crystalline powder, without odor or taste. It is very sparingly soluble in cold and hot water. It is a dibasic acid. The urates of the alkaline metals are more soluble than the acid. Lithium urate is more soluble than the potassium salt, and potassium more soluble than the sodium salt. The acid salts are less soluble than the normal, and the acid sodium salt frequently separates from the urine on cooling, or during **acid fermentation**, as a reddish-colored deposit—"brick-dust" or "lateritious" deposit.

Uric acid may be detected by evaporating the solution nearly to dryness, with nitric acid, when, on the addition of ammonia to the yellow residue, a deep reddish-violet color is produced (murexid). Under the microscope, uric acid from urine appears as yellow crystals, and is notable for the variety of forms it assumes.

ORGANIC BODIES CONTAINING NITROGEN.

AMINES, AMIDES, IMIDES, NITRILES.

601. Thus far we have said nothing of the **ammonia type** of compounds. Ammonia, $N\begin{cases}H\\H\\H\end{cases}$, is a type of a large number of compounds, most of which belong to organic chemistry.

An **amine** may be regarded as formed by replacing one or more atoms, in ammonia molecules, by positive or hydrocarbon radicals. When formed from one molecule of NH_3, they are monamines; when from two or three, they are called diamines or triamines.

Primary, secondary and tertiary amines are formed by replacing one, two, and three atoms of hydrogen respectively. We may represent these different classes of amines as follows:—

MONAMINES.

	Primary.	Secondary.	Tertiary.
$N\begin{cases}H\\H.\\H\end{cases}$	$N\begin{cases}C_2H_5\\H\quad .\\H\end{cases}$	$N\begin{cases}CH_3\\CH_3.\\H\end{cases}$	$N\begin{cases}CH_3\\C_6H_5.\\C_6H_5\end{cases}$
Ammonia.	Ethylamine.	Dimethylamine.	Methyl-diphenyl-amine,

DIAMINES.

$N_2\begin{cases}H_2\\H_2.\\H_2\end{cases}$	$N_2\begin{cases}C_2H_4\\H_2.\\H_2\end{cases}$	$N_2\begin{cases}C_6H_4\\C_6H_4.\\H_2\end{cases}$	$N_2\begin{cases}C_2H_4\\C_2H_4.\\(C_2H_5)_2\end{cases}$
Condensed Ammonia.	Ethylene-damine.	Diphenylene-diamine.	Diethylene-diethyl-diamine.

Amines all have basic properties and combine directly with acids, like NH_3, to form salts.

602. **Amides.**—These compounds may be regarded as formed by replacing the hydrogen atoms in ammonia with negative or acid radicals, as:—

$$\left.\begin{matrix}H\\H\\C_2H_3O\end{matrix}\right\}N.$$

Acetamide.

These bodies are classified and named like the amines. Secondary and tertiary amides, with very few exceptions, are unknown. They differ from the corresponding acids in containing amidogen, NH_2, in place of hydroxyl, OH. Thus:—

$HOC_2H_3O.$	$H_2O_2CO.$
Acetic Acid.	Carbonic Acid.
$H_2N—C_2H_3O.$	$(H_2N)_2CO.$
Acetamide.	Carbamide (urea).

Acid Amides are formed by replacing a part of the hydrogen in a polybasic acid by amidogen, NH_2, thus:—

$NH_2.HO.CO.$	$NH_2HOC_2O_2.$
Carbamic Acid.	Oxamic Acid.

The hydrogen of the amidogen in these compounds may be further replaced by hydrocarbon radicals, thus:—

$N_2H_3\begin{cases}C_2O_2\\C_6H_5.\end{cases}$	$NH\begin{cases}C_2H_2O—O—H\\C_7H_5O.\end{cases}$
Phenyl Oxamid.	Hippuric Acid.

603. **Imides.**—An imide differs from an amide in containing the radical NH, where the latter has NH_2, thus—

$$\underset{\text{Acid Carbamide}}{NH_2CO.OH} - H_2O = \underset{\text{Carbimide.}}{NHCO.}$$

The carbamides are converted by heat into isomeric bodies, which may be regarded as compounds of nitrogen with a triad radical, and known as **nitriles**, thus:—

$$\underset{\text{Diethyl Carbamide.}}{N_2\left\{\begin{matrix}(C_2H_5)_2\\H_2\\CO\end{matrix}\right.} \text{ gives } \underset{\text{Propionitrile.}}{N\left\{C_3H_5\right.}, H_2O \text{ and } \underset{\text{Ethyl Cyanide.}}{CN.C_2H_5.}$$

AMINES.

604. Of the large number of possible compounds which belong to this class of bodies, we can only mention one or two in detail. As before mentioned, the amines are bases, or organic alkalies. They unite directly with acids to form salts, and, with platinic and auric chlorides, to form double salts similar to those of ammonia. When heated, amines expel ammonia from its salts. The lower members of the group are gases, while the higher members are liquid or solid. The natural alkaloids seem to belong in this group, and will be mentioned in this connection.

As perhaps the most important artificial amine, we may mention **aniline** (phenyl amine), $N\left\{\begin{matrix}C_6H_5\\H.\\H\end{matrix}\right.$

605. **Aniline** is obtained from benzol, by first converting this into nitrobenzol by the action of concentrated nitric acid.—

$$C_6H_6 + HNO_3 = C_6H_5NO_2 + H_2O.$$

Nitrobenzol is generally a brown or pale yellow, strongly refractive liquid, boiling at 220° C. (428° F.); has a burning, sweet taste; an odor resembling that of bitter almonds and cinnamon. It is used in perfumery under the name of **essence of mirbane.** Taken internally, it is a violent poison. When treated with nascent hydrogen, generated by zinc and sulphuric or hydrochloric acid, **aniline** is produced.—

$$\underset{\text{Nitrobenzole}}{C_6H_5NO_2} + 3H_2 = \underset{\text{Aniline.}}{C_6H_5NH_2} + 2H_2O.$$

The aniline combines with the acid present, forming crystallizable salts. Like all the amines, it acts as a base, uniting directly with acids.

Aniline is, when pure, a colorless, oily, refractive liquid, boiling at 182° C. (359° F.), and is insoluble in water. It is the basis of a large number of beautiful coloring matters very much employed in dyeing and calico printing. In 1858 Perkins obtained a purple dye by acting upon aniline oil with potassium dichromate and sulphuric acid. Other colors were soon obtained; greens, yellows, reds, blues, violets and black have all been obtained from aniline. These dyes are soluble in alcohol and glycerine, and are used for a variety of purposes; they are not very permanent dyes.

Aniline is easily detected by the addition of a solution of sodium hypochlorite (chlorinated soda), with which it gives a purple-colored solution.

NATURAL ORGANIC BASES, OR ALKALOIDS.

606. Many plants, and especially those having medicinal and poisonous properties, contain basic principles, or compounds containing nitrogen, which are called **alkaloids.** Some are volatile, while others decompose when heated. Most of them resemble the amines or compound ammonias in properties, while some correspond more nearly to the ammonium compounds. Most alkaloids are sparingly soluble in water, more freely in alcohol, the solutions having an alkaline reaction and bitter taste. They combine directly with acids, like NH_3, forming crystalline salts, which are generally soluble in water. Their hydrochlorides form crystalline double salts with the chlorides of gold, platinum, mercury, etc. Most of them are precipitated by solutions of tannin, the double iodide of potassium and mercury, double iodide of cadmium and bismuth, by phosphomolybdic, phosphoantimonic, phosphotungstic, and picric acids, and by a solution of iodine in potassium iodide, or in hydriodic acid. These reagents are, therefore, used to separate the alkaloids from other substances found with them. By treating these precipitates with an alkaline hydrate, the bases are separated.

PROPERTIES OF NATURAL ALKALOIDS.

607. The following tables give the formulæ and solubility in water, benzol, ether and chloroform; also source, appearance and leading properties. The limits of this work will not allow us go further into this subject.

NATURAL ALKALOIDS.

NAME.	FORMULA.	SOLUBLE IN WATER.	AMMONIA WATER.	BENZOL.	ETHER.	CHLOROFORM.	SOURCE.
Aconitine	$C_{30}H_{47}NO_7$	S. in 150 pts.	S. S.	S.	S. in 2 pts.	S. in 2½ pts.	Aconitum Napellus.
Atropine	$C_{17}H_{23}NO_3$	S. in 60 pts.	S.	S. in 50 pts.	S. in 30 pts.	S. in 4 pts.	Atropa Belladonna.
Berberine	$C_{21}H_{18}NO_5 5Aq$	S. S.	S. S.	....	Insol.	S. S.	Berberis Vulgaris. Calumba root.
Brucine	$C_{23}H_{26}N_2O_4$	In 500 pts. boiling.	S.	S. in 60 pts.	Insol.	S. in 4 pts.	Nux Vomica.
Caffeine	$C_4H_5N_2O$	In 90 pts.	S.	S.	S. in 500 pts.	S. in 5 pts.	Coffee and Tea.
Cinchonine	$C_{20}H_{24}N_2O$	In 2500 pts. boiling.	Insol.	S.	S. in 400 pts.	S. in 60 pts.	Cinchona bark.
Cinchonidine	$C_{20}H_{24}N_2O$	In 2000 pts.	....	....	S. in 150 pts.	....	Cinchona bark.
Codeine	$C_{18}H_{21}NO_3$	In 75 pts.	In 75 pts.	S. in 12 pts.	S.	S.	Opium.
Colchicine	$C_{17}H_{19}NO_5$	Soluble.	S.	S. S.	S.	S.	Colchicum Autumnale.
Conine	$C_8H_{15}N$	In 100 pts.	S.	S.	S. in 6 pts.	S.	Water Hemlock. Conium Maculatum.
Delphinine	$C_{24}H_{35}NO_2$	Insol.	S.	S.	S.	S.	Delphinum Staphisagria.
Cocaine	$C_{17}H_{21}NO_4$	S. in 704 pts.	....	....	S.	...	Erythroxylon Coca.
Emetine	$C_{30}H_{44}N_2O_4$	S. S.	S. S.	S.	Nearly insol.	S.	Ipecacuanha.
Ergotine	$C_{50}H_{52}N_2O_3$	S.	....	...	Insol.	Insol.	Ergot of Rye.
Hydrastine	$C_{22}H_{23}NO_6$	Insol.	S. S.	...	S. S.	S.	Hydrastis Canadensis.
Hyoscyamine	$C_{15}H_{23}NO_3$	S. in hot.	S. in hot.	S.	S.	S.	Hyoscyamus Niger.
Morphine	$C_{17}H_{19}NO_3H_2O$	S. in 500 pts. boiling.	S. S.	Insol.	Insol.	S. in 90 pts.	Opium.
Narceine	$C_{23}H_{29}NO_9$	S. in 200 pts. boiling.	S. S.	S. S.	Insol.	S. S.	Opium.
Narcotine	$C_{22}H_{23}NO_7$	S. in 7000 boiling.	Insol.	S. in 25 pts.	S. in 120 pts.	S. in 3 pts.	Opium.
Nicotine	C_5H_7N	S.	S.	S.	S.	S. S.	Tobacco (Tabacum).
Papaverine	$C_{20}H_{21}NO_4$	Insol.	Insol.	S. in 40 pts.	S. S.	S.	Opium.

NATURAL ALKALOIDS.

NAME.	FORMULA.	SOLUBLE IN WATER.	AMMONIA WATER	BENZOL.	ETHER.	CHLOROFORM.	SOURCE.
[*] Physostigmine (Eserine).	$C_{15}H_{21}N_3O_2$	S. S.	S. S.	S.	S.	S.	Calabar bean.
Pilocarpine (Hydrochloride)	$C_{11}H_{16}N_2O_2HCl$	S. del.	S. S.	...	Insol.	Insol.	Jaborandi.
Piperine	$C_{17}H_{19}NO_3$	Nearly insol.	...	S.	S. in 90 pts.	S.	Cayenne pepper.
Pseudomorphine	$C_{17}H_{19}NO_4$	Insol.	Insol.	...	Insol.	Insol.	Opium.
Quinine	$C_{20}H_{24}N_2O_2$	S. in 1800 pts.	S.	S.	S.	S. in 50 pts.	Cinchona bark.
Quinidine	$C_{20}H_{24}N_2O_2$	S. in 750 pts.	...	S.	S. in 30 pts.	S.	Cinchona bark.
Rhœadine	$C_{21}H_{21}NO_6$	Insol.	Insol.	S. S.	S. in 1300 pts.	S. S.	Opium.
Strychnine	$C_{22}H_{24}N_2O_2$	S. in 6500 pts.	S. S.	S. in 160 pts.	Insol.	S. in 7 pts.	Nux vomica.
Thebaine	$C_{19}H_{21}NO_3$	Insol. (cold).	Insol.	S. in 18 pts.	S.	S. S.	Opium.
Theobromine	$C_7H_8N_4O_2$	S. in 750 pts.	S.	S. S.	Nearly insol.	S. S.	Cocoa Nuts.
Veratrine	$C_{32}H_{52}N_2O_8$	S. in 1000 pts. hot.	S. S.	S.	S. in 12 pts.	S. in 2 pts.	Veratrum Alba.
Urea	CH_4N_2O	S.	S.		S. S.		Urine.
GLUCOSIDE ALKALOIDS.							
Daphnin	$C_{31}H_{38}O_{19}$	S. S.	S.	...	Nearly insol.	...	Several species of Daphne.
Digitalin	$C_{10}H_{18}O_4$	S. S.	S.	S.	S. S.	S. S.	Foxglove. Digitalis.
Picrotoxin	$C_{12}H_{14}O_5$	S. in 50 pts.	S.	S.	S. in 250 pts.	S.	Cocculus Indicus.
Salicin	$C_{13}H_{18}O_7$	S.	...	Nearly insol.	Insol.	...	Willow leaves.
Solanin	$C_{43}H_{69}NO_{16}$	S. in 8000 hot.	Insol.	S. S.	S. in 4000 pts.	Insol.	Bittersweet.
Saponin	$C_{32}H_{54}O_{18}$	S.	S.	...	Insol.	...	Saponaria officinalis.

[*] Largely from Prescott's "Prox. Organic Anal."

EXPLANATION—S., soluble; S. S., slightly soluble. The numbers indicate the quantity of water required to dissolve one part of alkaloid.

608. **Separation and Detection.**—The alkaloids are separated from other substances and from each other by the use of various solvents, the principal ones being indicated in the preceding table. As most of the alkaloids or their salts are soluble in alcohol, this reagent is generally employed to extract them from plants or from other solid matters. After acidifying with acetic acid, digesting and filtering, the alcohol may be evaporated off, and the residue exhausted with ether, benzol, naphtha, chloroform, etc. Or, after acidifying, we may digest for some time with water, filter, and treat the filtrate with ammonium or sodium hydrate, or milk of lime. The alkaloids are thus precipitated, and may be separated by filtration, redissolved, converted into salts and purified by recrystallization.

The limits of this work will not permit us to enter into this subject fully, and the student is referred to works on chemical analysis for further particulars.

609. **Properties of Principal Alkaloids.**—**Aconitine.**—Glacial mass, or white powder, crystallizes with difficulty; in rhombic plates. It possesses a sharp, pungent taste, and in general, the physiological properties of the plant. With phosphomolybdic acid, it gives a yellow precipitate.

Aniline.—A colorless, oily liquid, turning dark on exposure; characteristic odor; boils at 182° C. (359° F.); gives red or violet color with NaOCl solution. (See Art. 605.)

Antipyrine, Dimethyl Oxichinine? (Alkaloid).—The hydrochloride is the salt used. It is a grayish-white, crystalline powder, with a weak, tarry odor, and bitter, somewhat acrid taste. It is readily soluble in alcohol and lukewarm water, from which it does not readily separate on cooling. Alkalies precipitate it, ferric chloride colors it reddish-brown, and nitrous acid colors it green, even in very dilute solutions (1 to 10,000). It is used, in doses of 15 to 30 grains, as an antipyretic.

Atropine.—Prisms; stellated tufts, white and fusible; bitter taste; dilates the pupil either when free or as salts. The alkaloid and its sulphate are officinal in the U. S. P.

Brucine.—Crystallizes in plates or prisms; lævogyrate; less poisonous than strychnine. With strong nitric acid gives a red, then yellow color; $SnCl_2$ gives deep violet.

Caffeine.—White, silky needles; fuses and sublimes; has a faint, bitter taste; poisonous in large doses; 0.4 to 0.5 grains produce death in cats and rabbits.

Cinchonine.—Four-sided needles, fusible. Its solutions are dextrogyrate; chlorine and ammonia give yellow precipitate;

separated from quinine by insolubility in ether; very bitter, and has less tonic properties than quinine. Sulphate officinal.

Cinchonidine.—Hard, rhombic prisms, with striated faces; differs from above in being lævogyrate. The sulphate is officinal.

Cocaine.—Monoclinic, glistening prisms, with bitter taste, leaving a benumbing effect on the tongue. Heated with strong HCl, it decomposes. The hydrochloride has powerful local anæsthetic properties. Coca leaves are used as a stimulant in Peru.

Codeine.—Rectangular prisms, isomorphous with morphine. Forms soluble crystalline salts, from which caustic soda precipitates the alkaloid.

Colchicine.—Yellow-white powder or colorless needles.

Conine.—Liquid; boils at 168° C. (335° F.); highly poisonous; odor of hemlock, and sharp, nauseous taste; decomposes on exposure to air.

Hyoscyamine.—Silky needles in groups, or amorphous and pasty; when impure, has a nauseous smell; dilates the pupil like atropine. It has sedative properties. The sulphate is officinal.

Morphine.—Short, transparent, trimetric prisms. When heated, loses water and melts to a colorless liquid; salts soluble in water; bitter taste; gives blue color with Fe_2Cl_6; decomposes iodic acid, giving free iodine. Noted for its anodyne properties; less exciting, and gives less after effects than opium. Officinal salts, sulphate, hydrochloride and acetate. When heated with large excess of HCl, it gives **apomorphine**, which is precipitated as a white powder by sodium carbonate, and turns green on exposure. Speedy emetic; said to be formed spontaneously in old morphine solutions.

OTHER OPIUM ALKALOIDS.

Oxymorphine (Pseudomorphine).
Codeine.
Thebaïne.
Codamine.
Laudanine.
Protopine.
Papaverine.
Meconidine.
Laudanosine.
Rheadine.
Cryptopine.
Narcotine.
Lanthopine.
Narceine.

Physostigmine (Eserine).—Crystalline or amorphous brown-yellow powder; solutions red to blue, of strong alkaline reaction; violent poison; strongly contracts the pupil. The salicylate is officinal.

Picrotoxine.—Colorless tufts of needles; reduces alkaline copper solutions; has an intensely bitter taste.

Piperine.—Colorless, tasteless, inodorous prisms, melting at 100° C. (212° F.). Alcoholic solution has sharp, peppery taste and neutral reaction. Soluble in alcohol.

Pilocarpine.—Uncrystallizable, but salts crystallize from alcohol. With H_2SO_4, it forms a colorless solution. The nitrate and hydrochloride are much used in medicine. Given internally, they produce rapid and profuse diaphoresis and salivation, quickened pulse and lowered temperature. With large doses, the heart stops in diastole. The hydrochloride is officinal.

Quinine.—Fine needles. Alkaloid seldom used. Sulphates, salicylate, tannate, hydrochloride and numerous salts in use. Forms basic, neutral and acid salts.

Basic Quinine Sulphate, $(C_{20}H_{24}N_2O_2)_2H_2SO_4.8H_2O$, is manufactured on a large scale, and is found as long, brilliant needles, which effloresce and form white powder. This salt is not very soluble (1 in 780 parts water; more readily in alcohol). All salts have an intensely bitter taste.

Normal Quinine Sulphate, $C_{20}H_{24}N_2O_2.(H_2SO_4)_2.7H_2O$, forms transparent four-sided prisms. Made from the basic salt by adding H_2SO_4. Soluble in eleven parts of cold water.

Acid Quinine Sulphate (Bisulphate), $C_{20}H_{24}N_2O_2(H_2SO_4)_2.7H_2O$, forms soft, white prisms, freely soluble in water, exhibiting beautiful blue fluorescence.

Solution of iodine precipitates from the acetic acid solution of basic sulphate, a tourmaline-like, crystalline body, of a green reflection.

The solutions of quinine and its salts have a left-handed rotary power on light. When dissolved in chlorine water, ammonium hydrate produces a deep emerald-green solution. If K_4FeCy_6 be first added, NH_4OH gives a deep red color. Quinine and its salts have a tonic and antiperiodic action on the economy. The sulphate, bisulphate, hydrobromide, hydrochloride and valerianate are officinal.

Quinidine.—Large prisms, soluble, with difficulty, in water, but soluble in alcohol. Gives quinine reaction with chlorine and ammonia; bitter taste and tonic effects, similar to quinine. Right-handed rotary power. Salts resemble those of quinine. The sulphate is officinal.

Salicin.—Tabular or scaly crystals. A glucoside. (See Art. 598.)

Strychnine.—Four-sided trimetric prisms, white and fusible. Bitter taste can be detected in a solution containing one part in one million parts of water. Forms soluble crystalline salts. Sulphate and phosphate are used largely in medicine. Violent poison, producing tetanic convulsions. The antidotes are morphine, atropine, and chloral hydrate. A delicate **test** is to dissolve in sulphuric acid and add a fragment of potassium dichromate, when a deep purple color is produced. The sulphate is officinal; but other salts are used in medicine.

Theobromine.—Small, trimetric crystals, in club-shaped groups. It has a slightly bitter taste. Salts are unstable, and decompose in contact with water.

Veratrine.—White prisms or powder, melting at 115° C. (239° F.) and solidifying, on cooling, to a resinous mass. Its dust causes violent sneezing; it is a violent poison. H_2SO_4 concentrated, forms a yellow solution, which gradually becomes red. It forms crystalline salts. The oleate is officinal.

PTOMAINES.

610. **Putrefactive, or Cadaveric Alkaloids.**—These alkaloids are produced during the putrid decomposition of animal and vegetable matter, and probably in certain pathological conditions in the human body during life. A distinct alkaloid has been separated from pyæmic fluid, and named **septicine.**

Some of these bases are very poisonous, producing symptoms resembling those caused by strychnine, atropine, conine, etc. Selmi obtained poisonous bases containing arsenic, from the body of a subject who had died of arsenical poisoning, and was exhumed fourteen days after death. It is probable that some of the symptoms of poisoning by preserved food, such as preserved fish, meat, etc., that are occasionally seen, are due to the presence of some one of these alkaloids. Tainted meat, fish, etc., should not be eaten. The author has known of a number of cases of poisoning from this cause.

Some of these alkaloids decompose with great ease, giving a cadaveric odor, while others remain permanent. Although they are most likely to be found in putrefying animal matters, they have been produced by the putrefaction of maize, leguminous substances, flour, etc. Many of the ptomaines are volatile and amorphous, but form crystalline salts with the acids. They answer to nearly all the ordinary reactions of the vegetable alkaloids. They seem to possess less stability, and have a greater tendency to undergo oxidation than the vegetable bases, and hence frequently act as reducing agents.

The number of the ptomaines seems to be large; and no one of them has been completely studied. The best test for the general detection of these bases, depends upon their power of reducing potassium ferricyanide to the ferrocyanide. The base is converted into the sulphate and brought into contact with a drop of potassium ferricyanide in a watch glass. A solution of ferric salt added to the watch glass will give a Prussian blue color if the ferricyanide has been reduced to the ferrocyanide, otherwise not. The reduction takes place rapidly. Unfortunately, a few of the vegetable alkaloids give this reaction as well as the ptomaines. No certain distinctive test has yet been devised.

GLUCOSIDES.

611. The glucosides are a large class of compounds widely distributed throughout the vegetable kingdom, which have the property of being resolved into a sugar and another compound, by acids, alkalies, or ferments. They are ethers of glucose. Some of them occur in the animal economy. We can mention here but a few of the more important of these bodies.

Amygdalin, $C_{20}H_{27}NO_{11}$, occurs in bitter almonds, in kernels of cherries, plums, apricots, and in leaves of laurel. Extracted from almonds by boiling alcohol, and precipitated by adding ether, it is obtained as pearly scales. Crystallizes in prisms. When **emulsin** (the ferment of bitter almonds) is added, it splits up into prussic acid, HCN, benzaldehyde, C_6H_5COH, and glucose.

Arbutin, $C_{13}H_{16}O_7$.—Extracted from leaves of Uva-Ursi. Soluble in water, has a bitter taste, and is crystalline. Emulsin and dilute acids yield glucose and hydroquinone.

Antiarin.—The active principle of the arrow-poison of Java; crystalline, soluble in water and alcohol. Obtained from the milky juice of **Antiaris toxicaria.**

Coniferin, $C_{16}H_{22}O_8$, occurs in cambial layer of the **coniferæ,** and crystallizes in stellate groups of prisms. Emulsin yields glucose and coniferyl alcohol. This latter, when treated with sulphuric acid and potassium dichromate, yields artificial **vanillin,** a body identical with that obtained from the vanilla bean. It is now manufactured on a considerable scale.

Convolvulin, $C_{31}H_{50}O_{16}$. Active principle of jalap; a resinous mass soluble in alcohol and alkalies. **Jalapin** exists with the above in jalap.

Digitalin.—A poisonous substance existing in common foxglove; forms an amorphous powder having an intensely bitter taste. Composition not known.

Esculin, $C_{21}H_{24}O_{13}$, and Esculetin, $C_9H_6O_4$, occur in the bark of the horse chestnut tree; sparingly soluble in cold, more freely in hot water, are crystalline and have a bitter taste.

Fraxin, $C_{32}H_{36}O_2$, is found in the bark of the ash and horse chestnut tree, and forms colorless needle-like crystals, soluble in water, furnishing a bitter, fluorescent solution.

Glycerrhicin, or **Liquorice-Sugar,** $C_{24}H_{36}O_9$, is the sweet principle of liquorice. It is a yellow, amorphous powder, having a sweet, acrid taste. It is soluble in water and alcohol.

Helleborin, $C_{36}H_{42}O_6$, is found together with **helleborein** in the root of green hellebore. It is insoluble in water, and forms glistening needles. A powerful poison.

Indican occurs in all plants yielding indigo, and sometimes in urine, and forms a pale-brown, syrupy liquid, having a bitter taste. The liver is probably the seat of its production in the body, and indol, taken up from the alimentary canal, the material from which it is produced. When allowed to ferment, or when treated with dilute acids, it forms indigo blue and indiglucin, a form of sugar. Indican is found in dogs' urine in greater quantity than in human urine. It also sometimes appears in the sweat. (See Indigotin, Art 625.)

Populin, $C_{20}H_{22}O_8$, occurs with salicin, in the bark and leaves of the aspen. It forms small prisms, having a sweet taste. Boiled with barium hydrate, it yields salicin and benzoic acid.

Phlorizin, $C_{21}H_{24}O_{10} + 2H_2O$, occurs in the root bark of the apple, plum, pear and cherry tree, and is soluble in alcohol. It is soluble in hot water, from which it crystallizes in silky needles having a bitter taste. Boiled with dilute acids, it yields glucose and **phloretin.**

Polychroite, $C_{48}H_{60}O_{18}$, is the coloring matter of saffron, and forms an amorphous, deliquescent, ruby-red mass.

Quercitin, or Flavin, $C_{27}H_{28}O_{12}$, occurs in tea, quercitron, sumac, grape vine, catechu, etc. It is slightly soluble in water, soluble in alcohol, and forms small, yellow crystals, which may be partially sublimed in beautiful, yellow needles. It is colored green by Fe_2Cl_6.

Saponin, $C_{32}H_{54}O_{18}$, occurs in **quillaia**, or soap-tree bark, and other plants. Soluble in water and alcohol. Its solutions behave like soap solutions. It is poisonous, but is sometimes added to soda water to produce a permanent froth. Its dust causes sneezing.

Solanin, $C_{42}H_{87}NO_{16}$, occurs in sprouted potatoes. It is a glucose alkaloid; soluble in alcohol, nearly insoluble in water, and forms gum-like salts.

Tannin, or more properly, tannic acids, form a group of bodies found in plants widely distributed. These bodies are soluble in water, have an acid reaction, an astringent taste, and form an insoluble compound with gelatin and albumen. They unite with animal skin, forming leather. With ferric salts, they form blue-black or green precipitates. They are used in the preparation of inks, in dyeing and tanning.

Tannic, or Gallotannic Acid occurs in oak bark, nutgalls, sumac and some other plants, in considerable quantities. It may be exhausted with a mixture of ether and alcohol. It is an amorphous, shining mass. Ferric salts give with it a bluish-black precipitate (ink), tartar emetic a white one. It precipitates starch, gelatin, albumin, and most alkaloids. Its watery solutions decompose when exposed to the air for some time, yielding gallic and ellagic acids. Dilute mineral acids, when boiled with it, give gallic acid and glucose.

Gallic Acid, $C_7H_6O_5$ (trioxybenzoic acid), occurs in the leaves of the wild vine, mangoes, and in small quantity in gall nuts, from which it can be exhausted with hot water. It may be obtained from tannic acid by boiling with dilute acids or alkalies, or by spontaneous fermentation.

Gallic acid crystallizes from water as fine, silky needles, readily soluble in three parts of boiling water, in ether and alcohol. It gives a blue-black precipitate with ferric salts, and reduces the salts of silver, mercury and gold. Its normal salts are permanent, but their solutions readily decompose. It has a strong astringent taste, and is used in medicine, in common with tannic, as a hæmostatic and astringent.

Quinic, or Quinotannic Acid, $C_7H_{12}O_6$, occurs chiefly in cinchona barks as a salt of quinine, but is also found in the bilberry and coffee bean. It occurs as oblique, rhombic prisms. It is soluble in water. On dry distillation it yields, among other products, benzoic acid and phenol.

Other tannins are known, as **caffeetannic,** of coffee; **quercotannic,** of oak; **catechutannic,** of catechu; **kinotannic,** of kino, etc., which vary in properties slightly, according to source.

NATURAL FATS AND FIXED OILS.

612. Almost all the fats and fixed oils are compound ethers of glyceryl, C_3H_5'''. They are found in both the animal and vegetable kingdoms. Some are liquid, while others are solid. Some oils remain permanent in the air, like olive oil, while others oxidize and thicken, like linseed and poppy oil. These latter are called siccative, or drying oils. The fats are insoluble in water, difficultly soluble in alcohol, but soluble in ether, petroleum naphtha, and carbon disulphide. The composition of natural oils has been partially considered while speaking of the fatty acids. So far as known, no fat consists purely of one substance, but of a mixture of oleate, palmitate and stearate of the triad radical, glyceryl.

These fats are decomposed by heat; acrolein being one of the products.

Stearine, or Stearin, $C_3H_5(C_{18}H_{35}O_2)_3$, is found in the more solid fats. It may be separated from the other principles, by melting tallow with turpentine, when the stearin remains in solution, while the olein and palmitin are precipitated.

By adding water to the liquid, the stearin may be separated. It fuses at 71° C. (160° F.), and solidifies again at 50° C. (122° F.).

Palmitin, $C_3H_5(C_{16}H_{31}O_2)_3$, is the chief constituent of mutton fat, lard and human fat. It is more soluble in alcohol and ether than stearin. It crystallizes in fine needles, and its melting point is variously stated at 46° C. (115° F.) and 61° to 62° C. (142° to 144° F.).

Olein, $C_3H_5(C_{18}H_{33}O_2)_3$, is the fluid constituent of most fats and oils. When pure, it is a colorless fluid, becoming yellow on exposure to the air. It may be obtained from olive oil by treating it with cold alcohol, cooling the solution to 0° C. (32° F.) to separate the palmitin, and adding water to the alcoholic solution to precipitate the olein. Olein is more abundant in vegetable than in animal oils.

When treated with hot alkalies or superheated steam, the fats are saponified (see Art. 553). Most fats decompose slowly in contact with air, and become rancid. In the process of digestion they are partially saponified and then emulsified; *i. e.*, broken up into minute drops. The active agents in this change are the bile and pancreatic secretion. The emulsification and absorption of partially saponified fats takes place with greater ease than with pure fat; hence, a slightly rancid oil is more easily assimilated than a fresh one.

The sources of fat in the human body are, 1st, the fat taken as food; 2d, the decomposition of proteids; 3d, the carbohydrates, a portion of which is converted into fat in the butyric fermentation in the intestines.

613. **Butter** consists of a mixture of stearin, palmitin and olein, not soluble in water, and the glycerides of butyric, caproic, caprylic and capric acids. These last acids are soluble and volatile. Oleomargarin, butterine, suine, etc., are artificial mixtures of butter with foreign fats, made to imitate butter.

The principal foreign fats employed are lard, beef oil, cotton seed, sesame and similar oils. The usual method of manufacture is to melt the foreign fats, deodorize them, when necessary, with nitric acid, then either to mix them with genuine butter, or churn them with milk. The mixture usually has a melting point above or below that of genuine butter.

Melting points of various fats :—

Butterine,	31.3° C.
Cocoa butter,	34.9° C.
Butter (average),	35.8° C.
Beef dripping,	43.8° C.
Veal dripping,	47.7° C.
Lard,	42° C. to 45° C.
Mutton fat,	50° C. to 51° C.
Tallow,	53° C.

A low melting point generally indicates butterine, or vegetable oils, while a high one indicates the presence of animal fats.

PROTEIDS.

ALBUMINOUS COMPOUNDS.

614. This important class of bodies forms the chief part of the solid constituents of blood, muscle, lymph, glands and other organs of animals, and is also found in plants, principally in the seeds. They are the principal substances taking part in the physiological changes in the organism. They are colloid (not crystalline), do not diffuse through animal membranes, and are very prone to putrefaction. No accurate formulæ have been found for them. They all contain carbon, hydrogen, nitrogen and oxygen, while most of them contain sulphur in addition, and all contain some ash, mostly in the form of calcium phosphate. Independent of the ash, they have the following composition:—

Carbon,	52 to 54	per cent.
Hydrogen,	7 to 7.3	"
Nitrogen,	13 to 16	"
Oxygen,	21 to 26	"
Sulphur,	1 to 1.6	"

The proteids are decomposed by heating with aqueous solutions of acids and alkalies, and with oxidizing agents.

When heated dry, they decompose, furnishing ammonia and compound ammonias. They all dissolve in strong acetic, hydrochloric and phosphoric acids, and in alkalies.

When boiled with hydrochloric acid, in contact with air, the solution assumes a blue color, changing to brown.

On heating with a solution of mercurous nitrate containing nitrous acid, they assume a fine red color, and the same color is produced by the joint action of a solution of sugar and sulphuric acid. When exposed to the air in a moist state, they rapidly putrefy, producing **ammonia**, ammonium sulphide, carbon dioxide, lactic, butyric, and other fatty acids, amines, leucin, tyrosin, etc.

ALBUMINS.

615. **Serum Albumin** exists in blood, chyle, lymph, and in small quantity in milk. In certain renal diseases, it appears in the urine. When its solutions are heated to about 72° C. (161.6° F.), it coagulates to a flocculent precipitate, which, when dried, forms a compact mass.

Its solutions have a specific rotary power of —56°. Strong mineral acids first precipitate it, then dissolve the coagulum. It is not precipitated by acetic acid alone, but when acidified with this acid, potassium ferrocyanide and ferricyanide coagulate it.

The following reagents may be used to detect its presence in solution; citric or acetic acid with potassium ferrocyanide, potassio-mercuric iodide, mercuric chloride, picric acid, concentrated nitric acid, or trichloracetic acid.

Egg Albumin is found in the white of eggs, and differs from the above in being almost insoluble in nitric and hydrochloric acids, and is precipitated by ether. Its specific rotary power is less than that of serum albumin, being —35.5°.

Vegetable Albumin occurs in small quantity in most vegetable juices. It shows the same general properties as the other albumins.

THE GLOBULINS.

616. The **globulins** differ from the albumins in being insoluble in water, but soluble in sodium chloride solution (1 per cent.). Except vitellin, they are precipitated by saturated solutions of the same salt. Soluble in very dilute HCl (1 part in 1000 being sufficient), with production of syntonin.

Vitellin occurs in yolk of egg. A white, granular body, soluble in dilute NaCl solutions, and not precipitated by a saturated solution of the same. It coagulates at about 70° C. (158° F.). Dissolves readily in dilute acid ($\frac{1}{10}$ per cent.) and in alkalies; it is precipitated by alcohol.

Globulin or Crystallin, occurs in blood, chyle, lymph, serous fluids, and the crystalline lens. In many of these fluids it exists in large quantity, and these coagulate spontaneously when removed from the living body, forming fibrin—probably by the combination of two modified forms of globulin called fibrinogen and fibrinoplastin. Globulin is not precipitated by a saturated solution of NaCl, but is precipitated by alcohol, or by carbonic anhydride allowed to bubble through the liquid.

Closely allied to globulin are two substances found in animal fluids; viz. **fibrinogen** and **fibrinoplastin or paraglobulin.**

Fibrinogen is found in blood, chyle, lymph, and in serous fluids, but especially in hydrocele and pericardial effusions.

Fibrinoplastin is met with in the serum of blood after coagulation, in white corpuscles, lymph, chyle, pus, connective tissue and the cornea.

These bodies are soluble in acetic acid, dilute alkalies, alkaline carbonates, and neutral saline solutions; but are precipitated from solutions by a current of carbon dioxide and very dilute acetic acid ($\frac{1}{10}$ per cent.) The neutral solutions of these bodies, when mixed in presence of a special ferment, form fibrin. This is supposed to explain the coagulation of blood, etc. Some authors claim that fibrinogen alone, with a soluble ferment, is concerned in the production of fibrin. These generators of fibrin are probably all derived from the white corpuscles.

Myosin, or muscle fibrin, is intermediate between **globulin** and **fibrin.** It is the name given to the solid which separates in the coagulation of muscle plasma, and forms an elastic gelatinous mass. Muscle juice, free from blood, is a yellowish, opalescent, syrupy, faintly alkaline liquid, coagulating at ordinary temperatures. The clot is called **myosin.**

Syntonin is easily prepared by treating myosin with a very small quantity of HCl.

Fibrin is a white, elastic, more or less fibrillated solid, insoluble in water or dilute salt solution. Soluble in acids (1 to 5 per cent.) with difficulty. With strong hydrochloric acid, it forms a violet solution. When boiled with caustic alkaline solutions, it forms ammonia and alkaline sulphides. It forms a blue color with hydric peroxide, or partially oxidized oil of turpentine, and freshly prepared tincture of guaiacum. Fibrin may be prepared by whipping blood with a bundle of twigs, and washing the coagulum with water and then with alcohol and ether.

HÆMOGLOBINS.

617. The **hæmoglobins**, or blood pigments, form the chief constituent of red blood corpuscles in vertebrates, and occur in the muscle of mammals, and in the blood of a few of the invertebrates. They all crystallize, but not with equal facility. All hæmoglobins are of a blood-red or brick-red color when in powder. They contain from .4 to .6 per cent. of iron, and differ slightly in composition.

The crystalline forms of hæmoglobins differ in different animals. The crystalline blood pigments are oxyhæmoglobins. Hæmoglobin forms a feeble compound with oxygen, which it releases by heating its solutions in a vacuum or in presence of ferrous sulphate, ammonium sulphide, stannous chloride, etc. Reduced hæmoglobin and oxyhæmoglobin are distinguished by their absorption spectra, the latter showing two such bands separated by a green band, while the former shows but one broad band occupying nearly the position occupied by the greenish-yellow band between the dark bands above mentioned. (See Figs. 10 and 11, Frontispiece). Hæmoglobin unites with nitric oxide, carbon monoxide, hydrochloric acid, etc. Some of these compounds give peculiar spectra; they are not easily expelled by oxygen, and hence are deadly poisons when inhaled.

Hæmatin, $C_{34}H_{34}N_4FeO_5$, is obtained in the form of a salt, by the decomposition of oxyhæmoglobin with an acid. It is an amorphous, blue-black mass, with a metallic lustre, insoluble in water or alcohol, but soluble in alkalies. It yields two different spectra; one with oxygen, and another with carbon dioxide.

Hæmatin Hydrochloride, or **hæmin** crystals, may be obtained by heating hæmoglobin or dried blood with common salt and glacial acetic acid. It forms thin, rhombic plates, having a brown-red color. The formation of these **blood crystals** is used to detect blood stains in criminal cases. The drop of dried blood is placed upon a microscopic slide, together with some pulverized common salt, and then treated with glacial acetic acid, and a cover glass placed upon it. Heat the slide until bubbles appear in the acid, cool, and examine with a ¼-inch objective. The presence of blood is indicated by the presence of the crystals.

ALBUMINATES (Derived Albumins).

618. There are two forms of these compounds, the **acid albuminate** and **alkali albuminate.** They are obtained by dissolving albumin in acids or alkalies. Some of the acid albumins contain sulphur, while the alkali albumins do not. When freshly prepared, they are soluble in dilute acids, alkalies, and alkaline carbonates, and their solutions are precipitated by careful neutralization, but not by boiling, and with difficulty by alcohol.

Alkali albumin occurs in blood cells and blood serum, chyle, pancreatic, nervous and corneal tissues, and in the crystalline lens. It closely resembles casein.

Acid Albumin, Syntonin, or Albumose.—Some authors limit the term **acid albumin** to the acetic acid compound, and syntonin to the hydrochloric acid compound, the first result of the action of gastric juice upon albumin. This distinction seems likely to confuse, and is, at least, unnecessary. Acid albumin, when pure, is a white, gelatinous mass, insoluble in water or salt solutions, but soluble in very dilute hydrochloric acid or caustic alkalies, and its solutions are not precipitated by boiling with water alone, but are coagulated by boiling with solutions of magnesium sulphate, strong salt solutions, alum, and many other metallic salts. Pure nitric acid gives a precipitate with its solutions.

Casein is a natural alkali albuminate found in milk of mammals, and differs only in some slight particulars from that found in blood, muscle, etc.

Casein exists in milk in the soluble form or in a state of semi-solubility. Under a high magnifying power the casein can usually be seen as a very finely granular solid, and therefore a portion of it, at least, must be in suspension, instead of in actual solution. There is a slight difference in the composition and properties of casein of cow's and of human milk, which explains a part of the difference in digestibility of the two. The former is more soluble in water and alcohol than the latter. Casein is somewhat richer in nitrogen than alkali albumin; it yields sulphur to heated potassium hydrate, which alkali albumin does not. It is coagulated by rennet, which is not the case with alkali albumin. Its solutions do not coagulate on boiling, are precipitated by most acids, but not by simple neutralization in presence of an alkaline phosphate. Potassium ferro-and ferri-cyanides, and dilute sulphuric acid precipitate casein in presence of free acetic acid. Rennet contains a special curdling ferment, as also do gastric and pancreatic juices, which can precipitate casein in alkaline solutions at slightly elevated temperatures.

Vegetable Caseins.—**Legumin** occurs in the leguminous plants, as an alkali albuminate. It may be obtained from softened peas, beans, etc. It resembles casein in its properties. It dissolves in dilute alkalies and acids, and when the albuminate solution is boiled, it forms a pellicle like milk. **Conglutin** resembles legumin, and is found in almonds, kernels of most stone fruits and lupins. **Gluten fibrin** occurs with other proteids in most of the cereal grains. Gluten, so called, is a mixture of several proteid substances. Gluten fibrin is insoluble in water and strong alcohol, but soluble in alcohol containing acetic or tartaric acid. Soluble in dilute acids and alkalies. Boiling water and salts coagulate it. **Mucedin,** and **gliadin,** are the gluten constituents soluble in alcohol.

PEPTONES.

619. The gastric, pancreatic, and probably the intestinal juices convert albuminous bodies into a more soluble and diffusible form, called **peptones.** Several different peptones have been described, but some of them are probably mixtures of partially peptonized albuminoids and true peptone; since the conversion is a gradual process of hydration, it is but natural to expect such a mixture. Dehydrating agents or simple heat can reverse the process, and convert peptone into albumin. It seems certain that the first action of gastric juice is to produce two bodies closely resembling, if not identical with syntonine (acid albumin). These are further changed into two kinds of peptone, named by Kühne **hemipeptone** and **antipeptone.**

The reactions of peptones are mostly negative. Solutions of peptone do not exhibit a viscid character, are not coagulated by heat, nitric acid, or acetic acid and potassium ferrocyanide. Alcohol precipitates it, but the precipitate is soluble in water. Tannin, mercuric chloride, picric acid, and potassio-mercuric iodide and acetic acid precipitate it.

The most marked property is its extreme solubility in water, and its ready diffusibility through animal membranes. Amyloid matter is not converted into peptone by pepsin or trypsin.

Lardacein, or Amyloid Matter is an amorphous, friable mass, occurring in certain regions of the body as a pathological product. It seems to be a derivative of fibrin. It is generally found as little transparent grains, or corpuscles, somewhat resembling starch granules. The most usual locations are the liver, spleen and kidneys. It gives many of the proteid reactions. It is stained a reddish-brown color with iodine, which is changed to a violet or blue tint by dilute sulphuric acid. Aniline violet stains it rose-red or violet instead of blue. Eosin stains it a bright red color.

THE COLLAGENS.

620. **Collagen** is the name given to the substance composing white, elastic tissue of the skin, tendons, etc. When boiled for some hours with water, it forms gelatin. **Collagen** and **ossein** seem to be closely allied in both composition and properties. Embryonic tissues when boiled, yield **mucin** and **chondrin** instead of gelatin. **Ossein** is the proteid basis of bones, and is converted into gelatin by boiling with water.

Gelatin in the pure state is a colorless or slightly yellowish, transparent, vitreous, tasteless mass. It swells in cold water, and readily dissolves in hot water or glycerine, forming a thick, viscid solution, which sets or gelatinizes on cooling. Heating to 140° C. (284° F.) or long continued boiling destroys this power of gelatinizing. It is soluble in dilute acetic and other acids, but insoluble in alcohol, ether and oils. Solutions of gelatin dissolve copper oxide with a blue color, which, on boiling, is reduced, but without separation of red oxide; it therefore interferes with the copper test for glucose. An impure gelatin or glutin prepared from animal refuse (bones, hides, etc.), when in the dry state, forms glue. Liquid glue is a solution of glue in acetic acid. Solutions of gelatin are precipitated by tannic acid, mercuric chloride, alcohol, and chlorine water, but not with acetic acid and potassium ferrocyanide, alum, or acetate of lead. Gelatin is lævogyratory.

Chondrin exists in permanent cartilages, and forms gelatin on boiling with water. The gelatin from this source differs slightly from that obtained from collagen, or ossein.

Mucin occurs in the cement substance of connective and epithelial tissues. It is also present in bile, and secretions of mucous membranes.

To prepare it from bile, precipitate with alcohol, dissolve in lime water, precipitate again with acetic acid, filter, and wash with alcohol or ether. It swells up in a little water, and dissolves in a large quantity. Its solutions do not coagulate on heating, and it contains no sulphur. It is insoluble in alcohol, ether, chloroform, or gastric juice.

Its solutions are precipitated by acetic acid, alum, basic acetate (sub-acetate) of lead, and very dilute mineral acids. Its solutions dissolve oxide of copper, and thus hinder the copper test for sugar.

BILIARY COMPOUNDS.

621. Bile contains a number of peculiar compounds, some of which occur in other animal secretions.

The three important classes of constituents are the acids, the biliary pigments, and cholesterin. In addition, we find lecithin, cholin, fats, soaps, mucus, a diastatic ferment, and inorganic salts. For analysis of human bile, see Appendix.

The bile of other animals contain some other ingredients not mentioned there, as **hyoglycocholic acid,** found in the bile of the pig, **lithofellic acid,** found in the bile of antelopes, **kyanuric acid,** found in the urine of dogs, and **lithuric acid,** found in calculi sometimes voided by oxen. **Glycocholic acid,** $C_{26}H_{43}NO_6$, occurs in bile as a sodium salt. The acid occurs in two forms, the one as fine crystalline needles, and the other as an amorphous resinous solid. It is monobasic. It is soluble in hydrochloric, sulphuric and acetic acids without decomposition. Soluble in glycerine, slightly soluble in cold and readily in hot water. Very soluble in alcohol and insoluble in ether. When boiled with alkalies or mineral acids, it splits up into **cholic acid** and **glycocin.**

Taurocholic acid, $C_{26}H_{45}NSO_7$, occurs in small quantity in human bile, but in larger quantity in that of the carnivora. It is soluble in water and alcohol. With boiling alkaline or acid solutions it forms **cholic acid** and **taurin.** The spontaneous decomposition of bile causes the same change.

Tests for Biliary Acids.—To a solution of the biliary acids add a few drops of a solution of cane sugar (1 to 10), and then strong sulphuric acid. A cherry red, followed by a deep purple violet color is produced. This test, known as Pettenkofer's, cannot be applied to organic mixtures, as urine, because numerous other bodies give the same color.

To apply it to such mixtures, evaporate to dryness, exhaust with absolute alcohol, decolorize with animal charcoal, evaporate to dryness, treat with water, and then test as above.

621. **Cholesterin,** $C_{26}H_{44}O$, may be obtained from the hot alcoholic solution of bile solids. It forms characteristic, glistening, rhombic plates, with notched edges or corners. It may be regarded as an alcohol, and forms a series of ethereal salts with

acids. It is soluble in benzol, chloroform, hot water, ether, and hot alcohol. Insoluble in cold water. Heated with nitric acid, it gives off yellow acid fumes and forms **cholesteric acid**, $C_8H_{10}O_5$.

Cholesterin forms the chief portion of biliary calculi; it is also found in blood, nervous tissue, the brain, yolk of eggs, and in the seeds of some plants.

Tests.—Cholesterin may be recognized: 1st, by the crystalline form under the microscope; 2d, when moistened with strong nitric acid, and evaporated to dryness, a yellow residue remains, which turns to a brick red on adding ammonia; 3d, on warming with dilute sulphuric acid, and then adding a drop of strong sulphuric acid it turns deep red in color.

The solution in chloroform, when shaken with an equal volume of strong sulphuric acid, forms a blood-red solution.

623. **Biliary Coloring Matters.**—There are at least four pigments obtainable from bile and biliary calculi, viz.: **bilirubin, biliverdin, bilifuscin** and **biliprasin. Bilihumin** and **hydrobilirubin** have also been described. The principal one of these, and probably the only one contained in bile when first secreted, is bilirubin, from which the others are derivatives.

Bilirubin, $C_{16}H_{18}N_2O_3$, is met, in the free state, in the bile of man and the carnivora, and ox bile; also, in combination with calcium, in biliary calculi.

It is probably identical with the red crystalline matter of old hemorrhagic clots, called **hematoidin,** and with **biliphaein, bilifulvin** and **cholepyrrhin.**

It may be prepared by treating the powdered biliary calculus first with ether, then with boiling water slightly acidified with hydrochloric acid. The residue is washed with pure water, and dissolved in hot chloroform and filtered; from the filtrate the chloroform is distilled off and the residue treated with absolute alcohol and ether, by which the bilifuscin is removed and bilirubin left. By dissolving in chloroform and setting aside, a part of it may be obtained as a dark red crystalline powder. It also exists as an orange red, amorphous powder.

It acts the part of a weak acid, combining with sodium, calcium, barium, lead, etc. Formula of the calcium salt, $(C_{16}H_{17}N_2O_3)_2Ca$.

Bilirubin is closely related to hæmatin, a derivative of hæmoglobin. This is its probable source.

$$3C_{32}H_{36}N_4FeO_6 + 3O = 6C_{16}H_{18}N_2O_3 + 3FeO \quad (?)$$

Moderately strong nitric acid, added to an ammoniacal solution of bilirubin, first colors it green, then blue (**bilicyanin**), violet red and finally yellow. (Gmelin's test for bile.) Nearly the same series of colors are produced by adding bromine to the chloroform solution, with the exception that it finally becomes colorless. Nascent hydrogen converts it into **hydrobilirubin**, $C_{32}H_{44}N_4O_7$, (**urobilin** of urine, **stercobilin** of fæces.) Hydrobilirubin is a dark brown amorphous powder, soluble in alkalies, sulphuric and acetic acids, alcohol, ether and chloroform. It does not give the play of colors with nitric acid. It is probably formed in the fæces by the action of the nascent hydrogen, set free by the butyric and putrefactive fermentations, upon bilirubin.

Biliverdin, $C_{16}H_{20}N_2O_5$ (or $C_{16}H_{18}N_2O_4$), is an oxidation product of bilirubin, and is readily formed by exposing alkaline solutions of this pigment to the air, or by similar treatment of fresh bile. It forms a green, amorphous powder, insoluble in water, ether and chloroform. It is soluble in alcohol, acetic acid and alkaline solutions, and reacts with nitric acid as does bilirubin. Hydrobilirubin may be prepared from it. **Biliprasin**, $C_{16}H_{22}N_2O_6$, is found in human gall stones, and **bilifuscin**, $C_{16}H_{20}N_2O_4$, occurs in small quantities in old bile, and in gall stones. According to Städeler, biliverdin is bilirubin + O and H_2O. Biliprasin is biliverdin + H_2O. Bilifuscin is bilirubin + H_2O. Viewed in this light these various pigments would appear to be formed in the following order: Hæmoglobin, hæmatin or hæmotoidin, bilirubin, hydrobilirubin (urobilin or urochrome and stercobilin), bilifuscin, biliverdin, biliprasin.

624. **Urinary Pigments.**—Besides urobilin, the urine contains at least one, and possibly more, coloring matters. Uroxanthin or indigogen, is a normal constituent of urine, but is much increased in the first stage of cholera, and in carcinoma of the liver. According to some authors, this coloring matter is identical with **indican.**

Uroxanthin may be detected in urine by adding to the urine its own volume of hydrochloric acid and a few drops of a solution of chloride of lime; the solution is colored red, violet, green, or blue, according to the amount of uroxanthin present. In some cases this test succeeds with hydrochloric acid alone.

In some cases the urine containing uroxanthin becomes blue, on standing for some days, from the spontaneous putrefaction.

Urobilin may be detected in fever urines by making them alkaline with ammonia, filtering, and adding a few drops of zinc chloride solution, when it will show a green fluorescence. This urobilin reaction may be obtained more distinctly by shaking the urine with ether, separating the ethereal solution, and, after evaporating the ether, dissolving the pigment in absolute alcohol. This solution will usually show the green fluorescence. **Urochrome** (Thudichum) **uromelanin** (Thudichum) **uroërythrin** and other pigments have been described, but the whole subject is enveloped in much uncertainty and confusion.

Black urine is occasionally seen after breathing arseniuretted hydrogen, in carbolic acid poisoning, and after inunctions of tar.

Melanin (Melanogen), is the black pigment of the choroid, melanotic tumors, and skin of the negro.

Pathologically, it is found in the urine of persons suffering with melanotic cancer, and sometimes with malaria. It is sometimes deposited in the lungs. Urine containing melanin turns dark on exposure to the air, or, more rapidly, with oxidizing agents, as nitric or chromic acids. Its detection is useful to the physician, as an aid to the diagnosis of melanotic cancer of liver, etc.

IMPORTANT VEGETABLE COLORING MATTERS.

625. **Indigo** is a blue coloring matter derived from several species of *Indigofera*, and other plants growing in India, Africa and South America. It exists as a glucoside, called **indican**, which is extracted with water; the liquid allowed to ferment in the air deposits the indigo as a blue powder. Commercial indigo is a mixture of several bodies, containing about 50 per cent. of indigo-blue or indigotin, $C_{16}H_{10}N_2O_2$.

Indigotin has a deep blue color, with a purple tinge; it is insoluble in water, alcohol, dilute acids and alkalies, but soluble in boiling aniline, and in Nordhausen sulphuric acid, forming **indigotinsulphonic** or **sulphopurpuric acid**, $C_{16}H_9N_2O_2SO_3H$. On neutralizing this solution with potassium or sodium carbonate, a blue precipitate of the potassium or sodium salt is obtained. These salts are insoluble in common salt solutions but soluble in pure water, and are met with in commerce under the name of "indigo carmine" or "indigo extract." The aqueous solution of these salts, is used as a test for glucose in urine. Indigotin sublimes when heated, the vapor condensing into blue crystals. Reducing agents convert indigo-blue to indigo-white, or hydro-indigotin, $C_{16}H_{12}N_2O_2$, which in the moist state oxidizes to indigo-blue. In dyeing with indigo, the goods are steeped in indigo-white, and then exposed to the air, when indigo blue is deposited in the cloth. Indigotin has been prepared synthetically.

Litmus is a purplish-blue coloring matter obtained from lichens; generally from *Lecanora tartarea,* by steeping in urine, adding lime and potassium carbonate. The mixture is exposed to the air for a few weeks, with frequent stirring, when a thick, blue solution is obtained. This solution is thickened up with plaster-of-Paris or chalk, formed into cakes, and cut into little cubes. The coloring matter of litmus is a weak acid, forming salts having a blue color, the commercial product being the potassium salt. With acids it becomes red, from the liberation of the acid. **Turmeric** is the root of *Cucurma longa*; it yields a yellow tincture, which turns brown with alkalies. It is used to a large extent to give a yellow color to various articles of food, as mustard, chow-chow, vermicelli, etc

Saffron is the stigmas of the flower of *Crocus sativa.* It yields to dilute alcohol a yellow coloring matter called **polychroit** or **crocin.**

Tinctura croci is officinal (U. S. P.).

Saffron is used to color certain articles of food, as butter, cheese, etc.

Annatto is a yellow color, obtained from the seeds of *Bixa orellana.* It is used in coloring butter, milk, cheese, etc.

Logwood, the wood of **Hematoxylon,** contains a purple dye, hematoxylin, used for purple and black dyes. It is officinal, and is used as a tonic and astringent. **Brazil wood** furnishes red dyes and lakes. **Cochineal,** the female of the insect *Coccus cacti* yields to boiling water and alcohol, a beautiful, red coloring matter, which, precipitated with alum and an alkaline carbonate, yields carmine. **Chlorophyll** is the name given to the green coloring matters of the leaves of plants. It occurs as microscopic granules distributed through the cell protoplasm in all the green portions of the plant. It may be extracted with alcohol, ether and benzol.

Very little is known of this body, but it seems to be composed of two coloring matters, a blue and a green. It contains iron, and is supposed to possess the power, under the influence of sunlight, of decomposing CO_2 and uniting the carbon to the elements of water. The yellow color of autumn leaves, is due to **xanthophyll,** an oxidation product of chlorophyll.

COMPOUNDS OF ANIMAL ORIGIN NOT OTHERWISE CLASSIFIED.

626. **Leucin,** $C_6H_{10}(NH_4)_2O_2$ (Amido-caproic acid) is found in various organs, intestinal contents, and sometimes in the urine, especially in yellow atrophy of the liver.

It may be obtained by pancreatic digestion of albuminous matters, or by treating horn shavings with dilute sulphuric acid. It is a white, crystalline solid, forming shining plates, soluble in water. It unites with both acids and bases.

Tyrosin, $C_9H_{11}NO_3$, is formed at the same time with leucin, and the two usually occur together. It crystallizes in fine needles, which aggregate into sheaf-like bundles or radiating spherical masses. It is but slightly soluble in cold water, and unites with both acids and bases.

Lecithin, $C_{44}H_{90}NPO_9$, is an amorphous, waxy substance, very hygroscopic, soluble in alcohol and ether. It is found in nerve tissue, especially the gray variety, in yolk of egg, semen, blood, milk, bile, etc. It plays the part of a fatty body. It is unstable, and is decomposed by boiling alkalies into glycerin-phosphoric, oleic, palmitic and stearic acids, neurin, etc.

Protagon, or **Cerebrote,** $C_{160}H_{308}N_5PO_{35}$ (Gamgee), forms the principal part of the white substance of Schwann. It is a glucoside, and lecithin is one of its decomposition products. It is insoluble, but swells up in water, and is soluble in warm alcohol, from which it crystallizes in fine needles.

Thudichum has recently described a number of other phosphorized substances obtainable from brain tissue, which he classifies under two heads as **kephalins** and **myelins.** Also a large number of other products, of which little is known.

Kreatin, $C_4H_9N_3O_2$, is a weak base, forming crystalline salts with several acids. It forms colorless, transparent prisms, sparingly soluble in cold and moderately soluble in hot water. It is found in muscles, Liebig's extract of meat, and in urine. It is undoubtedly an antecedent of urea.

Kreatinin is formed from kreatin by dehydration, and is usually found with the latter. It is a strong base.

Xanthin, $C_5H_4N_4O_2$, is met with in the spleen, muscles, pancreas, liver and urine. It seems to be an antecedent of uric acid. It forms white scales, sometimes depositing from urine as minute lemon-shaped plates. It is nearly insoluble in water, but soluble in dilute acids and alkalies. It is sometimes found in urinary calculi.

Hypoxanthin, $C_5H_4N_4O$ (sarkin), is a white crystalline powder, found in the spleen, thymus, muscular tissue, medulla of bones, urine, etc. **Allantoin**, **Carnin** and **Guanin** are allied to the last two bodies, all being compound ureas.

ALBUMINOID OR SOLUBLE FERMENTS.

627. The **soluble** or **unorganized ferments**, or **enzymes**, are a class of albuminoid bodies which have the power, under favorable circumstances, of causing certain chemical changes in other bodies with which they are brought in contact, without themselves undergoing any change. They are called ferments because of the similarity of their action to that of yeast and other well-known ferments. Some of these bodies are of vegetable, while others are of animal origin. Those of vegetable origin are **diastase**, **emulsin** and **myrosin**; while those of animal origin are **ptyalin** (salivary diastase), **pepsin**, **curdling ferment**, **pancreatic diastase**, **trypsin**, **invertin**, **emulsin ferment**, **histozym**, and probably others.

The exact chemical composition of these bodies is unknown, except that they are proteids. They are all soluble in water, are precipitated by alcohol and by lead acetate. They are very diffusible, lose their activity by being boiled with water, but are not precipitated. They have not yet been obtained in a state of absolute purity. Their action seems to be purely physiological, and not chemical. They impart nothing to or take nothing from the bodies they act upon.

Diastase, or **Maltin**, is the ferment formed from the gluten in the cereal grains at the time of sprouting Its chief object is the conversion of starch into dextrin and maltose. Ptyalin of saliva, and pancreatic diastase, if not identical with vegetable diastase, act in exactly the same way. They all act upon cooked starch with great rapidity, but have a very slow action upon unchanged starch. The process is one of hydration, and the action is similar to that which takes plåce when dilute sulphuric acid is boiled with starch or cellulose. When water is added to H_2SO_4 it probably forms H_6O_6S, ($H_2SO_4.2H_2O$) or orthosulphuric acid. This acid, when boiled, tends to part with a portion of its water, and if starch or other easily hydrated compound be present, it imparts this water to that body, in the **nascent state**, so to speak. Diastase acts upon the starch in the cold in the same way that H_6SO_6 does at a higher temperature. The first effect is to thoroughly liquefy the starch, then convert it into dextrin, and finally the dextrin into maltose.

The amount of starch that a given weight of diastase can thus transform is variously stated at from 2000 to 100,000 times its own weight, which, however, seems to be a fixed quantity with any given specimen of diastase. The rapidity of its action seems to depend upon the relative proportions of starch and ferment present. When the ferment is present in large quantity, the action is very rapid, almost instantaneous; while if it is small in proportion to the starch, it is slower in action. Diastatic ferment does not exist in the saliva and pancreatic juice of infants, previous to the sixth or seventh month, in sufficient quantity to digest much starch.

Emulsin or **Synaptase** occurs in sweet and bitter almonds. It may be extracted by digesting the almonds, freed from fat by pressure, for several hours with water. The filtered liquid is acidified with acetic acid, to precipitate conglutin, and the emulsin is then thrown down with alcohol, filtered off, washed with alcohol and dried. It is a white, friable mass, soluble in water, and capable of converting large quantities of amygdalin into sugar, prussic acid and benzoic aldehyde; it also converts salicin into sugar and saligenin. Its aqueous solution readily decomposes, yielding lactic acid.

Myrosin is the ferment of mustard.

Pepsin contained in gastric juice, is secreted by the glands of the stomach. It may be separated from the other constituents of filtered gastric juice by dialysis, as it does not diffuse through membranes. It is readily prepared by digesting the mucous membrane of the pyloric end of the stomach of the pig, first with strong alcohol, and after twenty-four hours expelling the alcohol by pressure, and digesting for some days with glycerin, slightly acidified with hydrochloric acid. Filter through muslin, then through paper, precipitate the pepsin with absolute alcohol, collect on a filter and dry. Other methods are in use.

Pepsin is a yellowish or grayish-white powder, soluble in water and glycerin, but insoluble in water. It gives none of the albumin reactions, and is precipitated by the acetates of lead. When dry, it may be heated to 110° C. (230° F.) without losing its activity, but its solutions lose it at a much lower temperature. Its activity is greatest at about 37° C. (98.6° F.), and requires hydrochloric, phosphoric, lactic or other dilute acid to develop its peculiar action. The presence of 0.1 per cent. of NaCl favors its action, but more than 0.5 per cent. hinders its action. Admixture of bile, carbolic acid, or excess of alcohol retard or entirely prevent its action. The acid of gastric juice is mostly hydrochloric during the intervals of digestion, but during digestion several organic acids are set free by the hydrochloric acid from the acetates, malates, tartrates, etc., taken with the food, so that the real work of the digestion is accomplished with the aid of various organic acids instead of hydrochloric alone. The specific action of pepsin is the change of proteids, whether coagulated or not, into syntonins and, finally, peptones. Pepsin is scarcely altered by putrefaction.

Trypsin occurs in the pancreatic juice, and may be extracted from the pancreas by the process described above for pepsin. Thus prepared, pancreatin is a yellowish-white amorphous powder, soluble in water and glycerin, but precipitated by alcohol. It possesses the property of acting upon the proteids in a way somewhat similar to pepsin, but is active only in alkaline solutions. It also converts starch into sugar, emulsifies and partially saponifies the fats, and curdles casein.

A part of the peptone at first formed by the pancreatic juice, is afterward converted into tyrosin and leucin. The digestion of the proteids is thus begun in the stomach, in an acid medium, and finished in the small intestine, in an alkaline medium. There is this difference in the two processes, that while acid pepsin readily liquefies the proteid bodies, it does not completely convert them into peptones; this completion of the process is more quickly and completely done by the trypsin.

Invertin is a ferment existing in the intestinal juices, which has the power of inverting cane sugar; *i. e.*, it converts it into dextrose and lævulose. Of its composition and other properties little is known. **Invertin**, or a substance possessing the same property, as well as a diastatic ferment, is found in the liquid portion of bakers' or brewers' yeast.

Histozym is a soluble ferment, supposed to exist in the blood, liver and kidneys, and which has the power of causing a variety of reactions within the body, such as the conversion of benzoic acid into hippuric acid, etc.

Some authors have described certain other ferments under the name of **microzimas**, which act as active chemical and physiological agents in the body during life, and cause its decomposition after death. They appear as minute molecular granulations, and are regarded as a part of the living organism by some, and as distinct organized ferments by others.

ORGANIZED FERMENTS.

628. Somewhat similar in action to the preceding group of ferments are certain forms of low vegetable organisms, which are known as **organized ferments.** These organisms vegetate most readily at temperatures of from 20° C. (68° F.) to about 40° C. (104° F.). Temperatures above or below these limits retard their growth, while a temperature near the boiling point entirely destroys their activity. A very minute quantity of any of these ferments can grow and exert its peculiar action as long as its peculiar nourishment lasts. Organized ferments excite chemical changes as the direct physiological result of their growth. They are all killed by hydrogen peroxide, and the chemical change is stopped by it. Borax stops the action of soluble or albuminoid ferments. All physiological fermentations in the organism are caused by the soluble ferments, while pathological fermentations are caused by the organized ferments.

The most important of the ferments of this class are yeast, or alcoholic ferment, acetic acid ferment, lactic, butyric and putrefactive ferments. With the exception of the first of these organisms, they all belong to the bacteria family.

Yeast (**Torula,** or **Saccharomyces cerevisiæ**) consists of one-celled, globular or oval-shaped plants, multiplying by budding. There are several varieties of this fungus. The principal action caused by yeast in saccharine fluids is, first, to convert the saccharose into grape sugar, and then change this into alcohol, carbon dioxide, and a trace of succinic acid and glycerin. The spores of yeast are always to be found either in the air or upon the surface of fruit, whence they find their way into the solutions made from their juices, which explains the apparent spontaneous fermentation.

Acetic acid ferment (**mycoderma aceti**) occurs usually in the form of chains of very small globular bodies, formed by multiplication of the cells by divisions, at right angles to the line of growth. It belongs to the bacteria family. It grows in alcoholic solutions containing a small amount of albuminous matter or ammoniacal salts, and alkaline and earthy phosphates. A little acetic acid favors its growth, as well as a free supply of air. It acts by causing an oxidation of the alcohol to acetic acid; when this change is complete, the ferment dies for want of nourishment.

Saccharomycetes Albicans (**Oidium Albicans**) is the ferment which is found growing upon the mucous membrane of the mouth of infants, producing the disease known as "thrush" or "sprue." The fungus appears as white patches upon the tongue and other parts of the mouth. The cells are globular, oval, or cylindrical, and occur in colonies or rows. It excites alcoholic fermentation, but feebly.

Lactic and **butyric** fermentations go hand in hand, the first usually, if not always, preceding the latter. They require a neutral or alkaline medium for their development, and grow best without oxygen, at a temperature of 35° C. to 40° C. (95° F. to 104° F.). These conditions exist in the intestines, and they are always found there. The subtances most prone to these fermentations are sugars, organic acids, soluble proteids, and especially mucus. The products of the fermentation are lactic, acetic and butyric acids, carbon dioxide and free hydrogen. These gases distend the bowel and often produce colic. Any excessive production of mucus in the bowel greatly favors these fermentations.

Putrefactive fermentation is caused by the growth of various forms of bacteria. The proteids are most liable to putrid fermentation, splitting up into fats, tyrosin, leucin, ammonia, sulphuretted hydrogen, carbon dioxide, hydrogen and nitrogen. This fermentation takes place in all organic infusions containing proteid matters, when exposed to the air; it also occurs in the small intestine in cases of constipation, or in some forms of indigestion, and to a slight extent in the normal condition.

POISONS AND THEIR ANTIDOTES.

629. Of the emergencies which arise in every-day life, or even in the practice of the young physician, none are more embarrassing than acute poisoning. The word poison, to many people, carries with it an idea of horror and panic. Even the physician is liable to something akin to alarm, when he comes into the presence of a victim of a violent poison. It is for this reason that we introduce a few of the most common poisons, with their antidotes. We do not intend to give a guide to the treatment of cases of poisoning, but merely a few simple rules to be remembered as first aids to those suffering with acute poisoning.

630. The first thing to be considered is, the **symptoms of poisoning.** Not unfrequently persons claim to have taken poison, when such is not the case. Or, suspicious friends fear that poison has been taken. The author has often met with cases of this kind, where a knowledge of the symptoms of poisoning has saved the patient a very disagreeable experience, and the

physician a great deal of trouble and future chagrin. The medical jurist should be familiar with the nature and actions of poisons, the symptoms which they produce, the circumstances which retard or otherwise modify their action their chemical and physiological antidotes, the pathological changes they induce, and the methods of combating these results. It is our purpose, here, to name a few of the symptoms of poisoning, and then to offer a few hints as to antidotal treatment.

The chief characteristics of poisoning are, more or less severe symptoms coming on suddenly, or within a few hours after taking some substance or fluid into the stomach, the individual being previously in a state of health. These symptoms usually increase steadily and uniformly, and tend to prove rapidly fatal. The symptoms may be greatly varied as to time and severity, by the quantity or form in which administered, the state of the stomach—whether full or empty—the condition of the person—whether asleep or awake—and a certain idiosyncracy of the individual. The symptoms which should arouse suspicion of acute poisoning are, the sudden onset of pain in the region of the stomach of a healthy person, especially of a "burning pain," accompanied by dryness of or a metallic taste in the throat, more or less vomiting, great prostration of the vital powers, a deathly or cadaveric aspect, or an expression of great fear or concern, the rapid intervention of coma, and speedy death. If all, or the greater number of the above symptoms are present in any case, there is reason for suspicion, and the physician should govern himself accordingly. (See remarks under Arsenic, Art. 244.) Poisons may, for convenience, be divided into the following five classes, based upon their effects upon the human subject: 1st, Corrosives; 2d, Irritants; 3d, Neurotics; 4th, Septic Poisons; 5th, Gaseous Poisons.

631. **Corrosive Poisons.**—To this class belong those poisons which exert, principally, a local action upon the tissues with which they come in contact. The most important of this class are corrosive sublimate ($HgCl_2$), the concentrated mineral acids (sulphuric, hydrochloric, nitric), and oxalic acid; the alkalies and their carbonates (potassium, sodium, and ammonium hydrates and carbonates); corrosive salts, as bisulphates of the alkaline metals, alum, nitrate of silver, chloride of zinc, butter of antimony ($SbCl_3$). Carbolic acid is a violent corrosive, when concentrated, and also has a remote effect upon the system after being absorbed. The symptoms of corrosive poisoning follow [illegible] after taking the poison, and are a sense of acid, alka-

line, or metallic, burning pain in the mouth, throat, gullet and stomach, usually inducing vomiting, which, however, does not relieve the distress. The pain soon extends over the entire abdomen, and is accompanied with symptoms of shock, or collapse.

632. **Irritant Poisons.**—Irritant poisons give rise to pain in the stomach, of a burning character, usually coming on some minutes or hours after taking the poison. In this respect, they differ in their action from the corrosives. The pain is accompanied, or followed, by vomiting, faintness, purging and tenesmus; the evacuations being often tinged with blood. The pulse is weak or irregular, and there is frequently severe headache. Death is usually caused by collapse, convulsions, or by inducing severe inflammations, which wear the patient out, after a variable period of time. Some have, also, a specific physiological action, besides their irritant action. The following are the more common irritants: Dilute mineral acids, concentrated organic acids, lime, zinc, copper, barium, silver and mercuric salts; all compounds of arsenic and antimony; phosphorus, iodine, bromine, etc. Many kinds of food may, under certain imperfectly understood conditions, become irritant poisons. Meat, fish, lobsters, tomatoes, etc., especially after having been canned and then exposed to the air.

633. **Neurotic Poisons.**—The neurotics exercise their action through the nervous system, and, therefore, only after absorption into the circulation. They rarely exert any local action. The neurotics are sometimes subdivided as follows:—

	Examples.
Narcotics, or those producing sleep	Opium.
Anæsthetics, or those producing insensibility	Chloroform.
Inebriants, or those producing intoxication	Alcohol.
Deliriants, or those producing delirium	Hyoscyamus.
Convulsives, or those producing spasms	Strychnine.
Hyposthenisants, or those producing death by syncope	Prussic acid.
Depressants, or those producing marked depression	Nicotine.

634. **Septic Poisons.**—To this class belong certain poisons introduced into the body through abrasions of the skin, open wounds, or by the fangs or sting of venomous animals or insects. In many respects these poisons resemble, in their action, the depressing narcotics.

635. **Poisonous Gases.**—To this class belong carbon monoxide (charcoal fumes), carbon dioxide (choke damp), marsh gas (fire damp), illuminating gas, hydrocarbon vapors, sewer gas,

confined air of living apartments, and noxious gases and vapors from manufacturing establishments.

TREATMENT OF ACUTE POISONING.

636. In every case of acute poisoning, or where the symptoms and circumstances indicate that a poison has been taken, **the first thing to do is to produce vomiting** (emesis), or to empty the stomach with the **stomach pump.** Then combat the local and constitutional effects of the poison by appropriate means.

THE PRINCIPAL EMETICS.

Zinc Sulphate.—Give 20 grains at once, or, dissolve ʒss in two ounces of water, and give tablespoonful every 15 minutes.

Copper Sulphate.—5 grains every 15 minutes, or, still better, 10 grains at once, followed by tepid water.

Alum.—A tablespoonful, given in syrup or honey.

Mustard (ground).—A dessertspoonful, stirred in tepid water and quickly swallowed. Very efficient, and is somewhat stimulating.

Apomorphine.—Give $\frac{1}{16}$ grain hypodermically.

Syr. Ipecacuanhæ.—Used mostly for children. Is depressant. fʒj every 15 minutes to a child two years old, until emesis is produced.

Tepid Water.—Drink copiously. It may be assisted by tickling the throat with a feather or the extended finger.

SPECIAL POISONS AND THEIR ANTIDOTES.

637. **Strong Mineral Acids.**—Chalk, lime, whitewash, magnesia, alkaline carbonates, baking soda, soap. Then give oil freely, and mucilaginous drinks.

Oxalic and **Tartaric Acids.**—Chalk, lime-water, whitewash.

After Treatment.—Mucilaginous drinks given freely, stimulants well diluted.

Alkalies, Caustic and Carbonated.—Vegetable acids, such as vinegar, lemon juice, tartaric or citric acids; fixed oils, such as castor, linseed, olive or cod-liver oil. Mucilaginous drinks may be given freely.

Antimony—Tartar Emetic—Wine of Antimony, or Oxide of Antimony.—Assist the vomiting by draughts of warm water, or mucilaginous drinks, such as flaxseed tea. Then give a cup of strong tea, or an infusion of oak bark, or a solution of tannin. This may be followed by opiates.

Chloride of Antimony.—Magnesia and baking soda.

Arsenic—White Arsenic—Arsenious Acid.—Expel the poison by thorough emesis. Promote the sickness by free use of albuminous or mucilaginous drinks. As an antidote, give raw eggs, beaten up in milk, freshly prepared ferric hydrate, solution of dialysed iron, or calcined magnesia, followed by stimulants well diluted.

Metallic Salts—Alum.—Alkaline carbonates, baking soda.

Soluble Barium Salts.—Soluble sulphates, Epsom or Glauber's salt.

Soluble Copper Salts.—Albumin, white of egg and milk, baking soda.

Iron—Green Vitriol—Persulphate of Iron.—Baking soda and emetics.

Lead—Sugar of Lead, White Lead.—Solution of Epsom or Glauber's salt, dilute solution of H_2SO_4.

Mercury—Corrosive Sublimate.—Albumin, white of egg and milk, followed by emetics.

Silver Nitrate—Lunar Caustic.—Common salt.

Zinc Chloride—Soldering Fluid, Burnett's Fluid.—Baking soda, milk, white of egg, tea, decoction of bark. Give opium to relieve the pain.

Tin, Chloride of.—Baking soda, magnesia, milk and white of egg.

Iodine—Most common form the tincture.—Give boiled starch paste, made thin enough to drink. In urgent cases, use starch or flour, with cold water.

Phosphorus—Rat Poison.—Has no true chemical antidote. Magnesia, milk of magnesia, chalk, or lime suspended in gruel, turpentine. Give no fixed oils.

Poisonous Meat, Fish, Lobsters, Etc.—Encourage vomiting by copious draughts of water; counteract depression with diluted brandy or whiskey; relieve pain with opium or one of its preparations. Apply hot fomentations to abdomen. When vomiting ceases, give castor oil or other laxative.

NEUROTICS.

Narcotics—Opium, Morphine, Laudanum, Paregoric, Soothing Syrups, Quieting Cordials, Etc.—First empty the stomach of any poison still remaining unabsorbed, by emetics. The patient is to be kept awake by forced walking, by the cold douche, or flagellations with wet towels. Faradaic electricity may be applied to the spine. Give strong coffee in abundance.

Anæsthetics—Vapors of Chloroform or Ether, Chloral, Methylene Dichloride, Nitrous Oxide, Etc.—Pure air, cold douches, artificial respiration, hypodermic injections of brandy, aquæ ammonia (diluted), or nitrite of amyl. Galvanism or Faradism may be employed, if the instruments are at hand.

Inebriants—Alcohol, Cocculus Indicus, Nitrobenzol (Essence of Mirbane), **Aniline, Etc.**—Emetics, when there is reason to believe that any poison remains unabsorbed, then ammonium carbonate, hydrate or acetate. Treat the narcosis as under opium.

Prussic Acid (hydrocyanic acid), **Potassium Cyanide, Laurel Water, Peach, Cherry, Plum Pits, Etc.**—No chemical antidote. Emetics, where there is time. Cold affusions to face and neck, inhalations of ammonia; spirits of ammonia should be given internally, with brandy.

Aconite.—See Prussic Acid.

Digitalis (Foxglove), **Calabar Bean** (Physostigma), **Tobacco** (Nicotine), **Hemlock** (Conium), **Lobelia Inflata** (Indian tobacco),	Emetics, stimulants, purgatives; stimulate with ammonia, alcohol or electricity.

Bites.—First wash thoroughly, then paint with carbolic acid (¼ strength), or tincture of iodine. Tie a handkerchief tightly above wound, until the doctor comes. Give alcohol freely in bites of snake, scorpion, tarantula, etc.

Stings.—Extract "stinger," if left behind. Apply mud, or a paste made of baking soda, or wash with ammonia.

Poisoned Wounds, **Dissecting Wounds,** **Infectious Diseases.**	Apply carbolic acid (¼ strength), or, paint with tincture of iodine. Give stimulants internally.

Poisonous Gases.—See special gases in text.

Treatment: Fresh air, rest, and mild stimulation.

APPENDIX.

We introduce here the following analyses of the various fluids of the body.

Subjoined is the **mean composition of the Blood**, as given by Becquerel and Rodier:—

	Per cent.
Water	78.16
Dry corpuscles	13.50
Albuminoids	7.00
Fibrin	0.25
Fats	0.17
Extractives	0.84
Earthy phosphates	0.08
Iron	0.05

The **ash** of human blood is thus given by Jarisch in the 100 parts:—

	Per cent.
Chlorine	30.74
Potash	26.55
Soda	24.11
Phosphoric acid	8.82
Sulphuric acid	7.11
Oxide of Iron	8.16
Lime and magnesia	1.33

HUMAN LYMPH, ETC.

	(Schmidt.) Per cent.	(Gubler and Quevenne.) Per cent.	(Hensen and Dähnhardt.) Per cent.
Water	96.39	93.48	98.52
Solids	3.6	6.52	1.48
Fibrin	—	0.6	
Globulin substances and serum albumin.	2.9	4.28	0.68
Fats, cholesterin, lecithin.		0.91	
Extractives.		0.43	
Sugar.		0.05	
Salts	—	0.82	0.79
Sodic chloride	0.54		
Soda	0.15		

ASH OF HUMAN LYMPH (Dähnhardt, in 100 parts of Ash).

	Per cent.
Sodic chloride	74.84
Soda	10.35
Potash	3.22
Lime	0.93
Magnesia	0.26
Phosphoric acid	1.09
Carbonic acid	8.21
Sulphuric acid	1.27
Ferric chloride	0.06

HUMAN CHYLE (O. Rees).

	Per cent
Water	90.48
Solids	9.52

Fibrin	a trace
Albumin	7.08
Fats, lecithin, cholesterin, etc	0.92
Extractives	1.0
Salts	0.44

In the ether extract of chyle obtained from a human fistula, the following bodies were present in the 100 parts:—

	Per cent.	Per cent
Cholesterin	11.3	to 14.1
Lecithin	7.5	" 8.8
Olein	38.1	" 77.1
Palmitin and Stearin	43.0	

MIXED SALIVA—HUMAN.

	(Jacubowitsch.) Per cent.	(Hammerbacher.) Per cent.
Water	99.51	92.42
Solids	0.48	0.58
Soluble organic bodies (ptyalin, etc.)	0.13	0.14
Epithelium	0.16	0.22
Inorganic salts	0.182	0.22
Potassic sulphocyanate	0.006	0.004
Potassic and sodic chlorides	0.084	——

ANALYSIS OF GASTRIC JUICE (HUMAN) MIXED WITH SOME SALIVA.

(After C. Schmidt, etc.)

	Per cent.
Water	99.44
Solids	0.56
Organic substances (pepsin and peptones)	0.32
Free hydrochloric acid	0.25
Sodic chloride	0.14
Potassic chloride	0.05
Calcic chloride	0.006
Phosphates of lime, magnesia, and iron	0.015

ANALYSIS OF PANCREATIC JUICE—DOG.

(*C. Schmidt.*)

	From a freshly opened pancreatic duct. Per cent.
Water	90.07
Solids	9.93
Organic substances as albumin, alkali albumin	9.04
Inorganic salts	0.89
Soda combined with albumin, and traces of lime and magnesia similarly combined	0.09
Sodium chloride	0.73
Potassium chloride	0.02
Phosphates of lime, magnesia, and soda	0.05

In **post-mortem** BILE, Hoppe Seyler obtained :—

	Per cent.
Glycocholate of soda	3.03
Taurocholate of soda	0.87
" " " (containing sulphur)	0.05
Soaps	1.39
Mucin	1.29
Lecithin	0.53
Cholesterin	0.35
Other organic substances insoluble in alcohol	0.14
Iron, probably as phosphate	0.006

In 100 parts, dry solids of bile **obtained from a fistula,** having a specific gravity of 1010, and 2.24 to 2.28 per cent. of solids, Jackson obtained :—

	Per cent.
Glycocholate of soda	44.8
Sodic chloride	24.5
Palmitate and stearate of soda	6.4
Phosphate of soda	5.9
Carbonate " "	4.2
Cholesterin	2.5
Phosphate of lime	1.6
Potassic chloride	1.2
Lecithin	0.2
Fats	0.4
Residue insoluble in alcohol and ether	8.1

ANALYSIS OF HUMAN MILK AND COW'S MILK (König).

	Woman's Milk.			*Cow's Milk.*		
	Mean.	Minimum.	Maximum.	Mean.	Minimum.	Maximum.
Water	87.09	83.69	90.90	87.41	80.32	91.50
Total solids	12.91	9.10	16.31	11.59	8.50	19.68
Fat	3.90	1.71	7.60	3.66	1.15	7.09
Milk sugar	6.04	4.11	7.80	4.92	3.20	5.67
Casein	0.63	0.18	1.90	3.01	1.17	7.40
Albumin	1.31	0.39	2.35	0.75	0.21	5.04
Albuminoids	1.94	0.57	4.25	3.76	1.38	12.44
Ash	0.49	0.14	?	0.70	0.50	0.78

ANALYSIS OF HUMAN TEARS (Lerch).

This fluid has the following simple composition:—

	Per cent.
Water	98.00
Albumin, with traces of mucus	0.50
Sodium chloride	1.40
Other salts, as alkaline and earthy phosphates	0.02

SWEAT.

The following is the composition of human sweat, according to Picard:—

	Per cent.
Water	98.88
Solids	1.12
Salts	0.57
Sodium chloride	0.22 to 0.83
Alkaline sulphates, phosphates, lactates, and potassium chloride	0.18
Fats, fatty acids and cholesterine	0.41
Epithelium	0.17
Urea	0.08

NORMAL URINE OF THE TWENTY-FOUR HOURS.

(After Liebermann, as quoted by Charles).

Constituents.	Amt. in grains.	Alterations Under Pathological Conditions.
Water.	50 oz.	**Increased** in diabetes, after absorption of effusions, in contracted kidney, etc. **Diminished** in acute fevers, cholera, dropsies, etc.
Sp. gravity.	1.020	**Raised** in diabetes mellitus, and occasionally in diabetes insipidus. **Lowered** in polyuria and certain cachectic conditions, depending on want of food etc.
Total solids.	900 to 1100	Fluid ounces × by last two figs. of sp. gr. = total solids.
Urea.	400 to 600	**Increased** in fevers to the crisis, in intermittent fever before the cold stage, in diabetes, and after absorption of dropsical effusions. **Diminished** in dropsies, in chronic liver diseases, in Bright's disease, after fevers, and in all conditions in which tissue change is hindered.
Uric acid.	5 to 12	**Increased** in acute fevers, in diseases of the lungs interfering with respiration (as tubercular deposit, etc.), acute rheumatism, leukæmia. **Diminished** in diabetes, chronic gout, Addison's disease.

Normal Urine of the Twenty-Four Hours. (Continued.)

Constituents.	Amt. in grains.	Alterations Under Pathological Conditions.
Kreatinin.	10 to 18	**Increased** in acute fevers, pneumonia, etc. **Diminished** in diabetes mellitus, debility, kidney disease, etc.
Hippuric acid	6 to 15	**Increased** in fevers, diabetes mellitus, and chorea.
Sulphuric acid	23 to 38	Having more or less the same source as urea, it will increase or diminish therewith.
Phosphoric acid	48 to 54	**Increased** in fevers, in most acute nerve affections, and in tubercle of the lung. **Diminished** in many mental diseases, especially mania, and in chlorosis.
Oxalic acid	0.3	**Increased** in catarrhal jaundice, and in oxalic acid diathesis.
Carbolic acid.	0.015	**Increased** in certain diseases of the intestines, causing constipation (ileus, etc.), but has been observed to be increased also in certain cases of diarrhœa.
Phosphate of lime.	4 to 5	**Increased** in osteomalacia, rickets, scrofula, carcinoma, long continued suppuration. **Diminished** in fevers.
Phosphate of magnesium.	7 to 11	
Chloride of sodium.	150 to 200 Cl = 90 to 120 Na = 60 to 80	**Increased** in fevers at the outset, and with the re-absorption of dropsical fluids. **Diminished** during apyrexia, dropsies, cholera, typhus, inflammations generally, and especially pneumonia.
Free acid (calculated as oxalic acid).	30 to 60	**Increased** during the acme of acute febrile affections (on account, probably, of the diminished proportion of water present). **Diminished** in most diseases affecting the nutrition and leading to a deficiency therein.
Indican.	0.07 to 0.05	**Increased** with diseases attended by constipation, and occasionally, also, in cases of diarrhœa.
Total inorganic salts.	200 to 380	
Potassium.	38 to 48	
Sodium.	140 to 180	
Calcium.	4 to 5	
Magnesium.	2 to 3	

TABLE OF WEIGHTS AND MEASURES.

ENGLISH WEIGHTS.

TROY WEIGHT.

Pound.	*Ounces.*	*Pennyweights.*	*Grains.*		*French Grammes.*
1	12	240	5760	=	373.2419
	1	20	480	=	31.1035
		1	24	=	1.5552

APOTHECARIES' WEIGHT.

℔. *Pound.*	℥ *Ounces.*	ʒ *Drachms.*	℈ *Scruples.*	gr. *Grains.*		*French Grammes.*
1	12	96	288	5760	=	373.2419
	1	8	24	480	=	31.1035
		1	3	60	=	3.8879
			1	20	=	1.2959
				1	=	.0648

AVOIRDUPOIS WEIGHT.

Pound.	*Ounces.*	*Drachms.*	*Grains.*		*French Grammes.*
1	16	256	7000	=	453.5926
	1	16	437.5	=	28.3495
		1	27.343	=	1.7718

METRIC MEASURES.

MEASURES OF LENGTH.

1 Millimetre = 0.001 of a metre.
1 Centimetre = 0.010 of a metre.
1 Decimetre = 0.100 of a metre = about 4 inches.
1 Metre = 1.000 Metre = 39.37 inches.
1 Decametre = 10.000 metres.
1 Hectometre = 100.000 metres.
1 Kilometre = 1 000 000 metres = about ⅝ of a mile.
1 Myriametre = 10,000.000 metres = about 6⅓ miles.

MEASURES OF SURFACE.

1 Centiare = 1 Square metre = about 1⅕ square yards.
1 Are = 100 Square metres.
1 Hectare = 10,000 Square metres = about 2½ acres.

MEASURES OF VOLUME.

1	Cubic centimetre	=	0.001 of a litre.
1	Litre (cubic decimetre)	=	1000. cubic centimetres.
1	Cubic metre	=	1000. cubic decimetres.
1	Cubic metre	=	1000. litres, or 1 kilolitre.
1	Cubic metre	=	1 stere.

MEASURES OF WEIGHT.

1	Milligramme	=	0.001 of a gramme	=	about $\frac{1}{65}$ of a grain.
1	Centigramme	=	0.010 of a gramme.		
1	Decigramme	=	0.100 of a gramme.		
1	**Gramme***	=	**1.000 Gramme**	=	**about 15½ grains.**
1	Decagramme	=	10.000 grammes.		
1	Hectogramme	=	100.000 grammes.		
1	Kilo(gramme)	=	1000.000 grammes	=	about $2\frac{1}{5}$ ℔s.
1	Tonneau	=	1000. Kilo's	=	about 1 ton.

ALPHABETICAL TABLE OF EQUIVALENT MEASURES.

1	Are = 100 Sq. metres,		119.6 Sq. yards.
1	Centimetre = $\frac{1}{100}$ metre,		.3937 inches.
1	Cubic centimetre (of dist. water weighs 1 gramme)		.061 cub. in.
1	Cubic decimetre (1 litre),	=	Imperial measure, 1.133 qts.
1	Cubic decimetre,	=	Wine measure, 0.8806 qts.
1	Cubic decimetre of water weighs 2.2046 ℔s. Av.	=	1000 grammes.
1	Cubic foot = 1728 cu. in.	=	28315.31 c.c.
1	Cubic foot of water weighs at 62° F. (16.6° C.),		62.32 ℔s. Av.
1	Cubic inch,		16.386 c.c.
1	Cubic inch of water weighs at 16.6° C. (62° F.),		252.46 grains.
1	Cubic inch of water weighs at 16.6° C. (62° F.),		16 372 grammes.
1	Cubic metre (1 stere) = 1000 litres	=	35.30 cu. ft.
1	Drachm, troy, = 3.888 grammes = ʒj.		
1	Fluid ounce, imperial, = 28.4 c.c.	=	1.7329 cu. in.
1	Fluid ounce, wine measure, (f℥j) = 29.57 c.c.	=	1.8047 cu. in.
1	Fluid ounce water, at 16.6° C. (62° F.), wine measure, weighs		456. grains.
1	Fluid ounce water, at 15.6° C. (60° F.), wine measure, weighs		29.57 grms.
1	Fluid drachm (fʒj) = 3.697 c.c.		
1	Foot, 12 in.,		30.48 centimetres.
1	Gallon, imperial, = 277.27 cu. in.	=	4.543 litres.
1	Gallon, wine, = 231 cu. in.	=	3.785 litres.
1	Grain, troy,	=	0.06479895 grams
1	Gramme (weight of 1 c.c. of water) at 4° C. (39.2° F.)	=	15.4323487 grains.
1	Inch	=	2.54 centimetres.
1	Kilogramme (1000 grms.)	=	2.2046 ℔s. Av.
1	Litre (see cubic decimetre),		61.027 cu. in.
1	Metre (one-forty millionth of earth's meridian),		39.3708 inches.
1	Minim = 0.0616 c.c. of water weighs 0.0616 grammes,	=	0.95 gr.
1	Ounce, troy, = 480 grs.	=	31.1 grammes.
1	Ounce, Avoirdupois, = 437.5 grs.	=	28.35 grammes.
1	Pint, imperial, = 567.93 cu. in., wine measure,	=	473.15 cu. in.
1	Pound, troy, = 5760 grs.	=	373.24 grammes.
1	Pound, Avoirdupois, = 7000 grs.	=	453.59 grammes.
1	Ton, Avoirdupois, = 2000 ℔s.		907.2 kilo's.
1	Tonneau, 1,000,000 grammes.		1000 kilo's.

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* Sometimes spelled *gram* in English and American books.

TABLE OF SPECIFIC GRAVITIES NAMED IN THE U. S. PHARMACOPŒIA, 1880.

(Compiled by Prof. P. W. Bedford.)

Acid Acetic	1.048
" " dil	1.0083
" " Glacial	1.056–1.058
" Hydrobromic dil	1.077
" Hydrochloric	1.160
" " dil	1.049
" Lactic	1.212
" Nitric	1.420
" " dil	1.059
" Oleic	.800–.810
" Phosphoric	1.347
" " dil	1.057
" Sulphuric	1.840
" " Aromat	.955
" " dil	1.067
" Sulphurous	1.022–1.023
Æther	.750
" Acetic	.889–.897
" Fortior	.725
Alcohol	.820
" dil	.928
Amyl Nitras	.872–.874
Aq. Ammon	.959
" " Fort	.900
Bals. Peru	1.135–1.150
Benzinum	.670–.675
Bromum	2.990
Camphora	.990–.995
Carbonei Bisulphidum	1.272
Cera Alba	0.965–0.975
Cera Flava	.955–.967
Cetaceum	0.945
Chloroform Puris	1.485–1.490
" Venale	1.470
Copaiba	.940–.993
Creasote	1.035–1.085
Fel Bovis	1.018–1.028
Glycerinum	1.250
Hydrargyrum	13.5
Iodoformum	2.00
Liq. Ammon. Acet	1.022
" Calcis	1.0015
" Ferri Acetatis	1.160
" " Chloridi	1.405
" " Citratis	1.260
" " Nitratis	1.050
" " Subsulph	1.555
" " Tersulph	1.320
Liq. Hydrarg. Nit	2.100
" Plumbi Subacetatis	1.228
" Potassæ	1.036
" Potassii Citratis	1.059
" Sodæ	1.059
" " Chloratæ	1.044
" Sodii Silicatis	1.300–1.400
" Zinci Chlor	1.555
Mel	1.101–1.105
Oleum Adipis	0.900–0.920
" Ætherum	0.910
" Amygd. Amar.	1.060–1.070 1.043–1.049
" " Express	.914–.920
" Anisi	.976–.990
" Aurantii Cort	.890
" " Flor	.850–.890
" Bergamii	.860–.890
" Cajuputi	.920
" Cari	.920
" Caryoph	1.050
" Chenapodii	.920
" Cinnamomi (*Ceylon*)	1.040
" " (*Chinese*)	1.060
" Copaiba	.890
" Coriandri	.870
" Cubebæ	.920
" Erigerontis	.850
" Eucalypti	.900
" Fœniculi	.960
" Gaultheriæ	1.180
" Gossypii Sem	.920–.930
" Hedeomæ	.940
" Juniperi	.870
" Lavendulæ	.890
" " Flor	.890
" Limonis	.850
" Lini	.936
" Menth. Pip	.900
" " Virid	.900
" Morrhuæ	.920–.925
" Myrciæ	1.040
" Myristicæ	.930
" Olivæ	.915–.918
" Picis Liquida	.970
" Pimentæ	1.040
" Ricini	.950–.970
" Rosæ	.860

Table of Specific Gravities. (*Continued.*)

Oleum Rosmarini	.900	Resina	1.070–1.080
“ Rutæ	.880	Sp. Ætheris Nitrosi	.823–.825
“ Sabinæ	.910	“ Ammoniæ	.810
“ Santali	.945	“ Ammoniæ Aromat	.885
“ Sassafras	1.090	“ Frumenti	.930–.917
“ Sesami	.914–.923	“ Vini Gallici	.941–.925
“ Sinapis Volatile	1.017–1.021	Syrupus	1.310
“ Succini	.920	Syr. Acidi Hydriodici	1.300
“ Terebinthiniæ	.855–.870	Thymol	1.028
“ Thymi	.880	Tinct. Ferri Acetas	0.950
“ Tiglii	.940–.955	“ “ Chloridi	0.980
“ Valerianæ	.950	Vinum Album	.990–1.010
Petrolatum	.835–.860	“ Rubrum	989–1.010
Phosphorus (at 50° F.)	1.83	Zincum	6.9

See List of Tests by Hans Wilder, Appendix.

Table showing the Solubility of some Chemicals in Glycerine. One Hundred parts of Glycerine Dissolve the Annexed Quantities of the Salts.—(Klever).

	Parts		Parts
Arsenious oxide	20.00	Morphine acetate	20.00
Arsenic oxide	20.00	“ chlorhydrate	20.00
Acid, benzoic	10.00	Phosphorus	0.20
“ oxalic	15.00	Plumbic Acetate	20.00
“ tannic	50 00	Potassium arseniate	50.00
Alum	40.00	“ chlorate	3.50
Ammonium carbonate	20.00	“ bromide	25.00
“ chloride	20.00	“ cyanide	32.00
Antimony and potassium tartrate	5.50	“ iodide	40.00
Atropia	3.00	Quinine	0.50
“ sulphate	33.00	“ tannate	0.25
Barium chloride	10.00	Sodium arseniate	50.00
Brucia	2.25	“ bicarbonate	8.00
Cinchonine	0.50	“ borate	60.00
“ sulphate	6.70	“ carbonate	98.00
Copper acetate	10.00	“ chlorate	20.00
“ sulphate	30.00	Sulphur	0.10
Iron and potassium tartrate	8.00	Strychnine	0.25
“ lactate	16.00	“ nitrate	4.00
“ sulphate	25.00	“ sulphate	22.50
Mercuric chloride	7.50	Urea	50.00
Mercurous chloride	27.00	Veratrine	1.00
Iodine	1.90	Zinc chloride	50.00
Morphine	0.45	“ iodide	40.00
		“ sulphate	35.00

SOLUBILITY OF THE MOST IMPORTANT CHEMICALS USED IN MEDICINE, IN WATER AND ALCOHOL.

(Columns 1 and 3 taken from U. S. Pharmacopœia).

Explanation of Signs.

s. soluble; *ins.* = insoluble; *sp.* = sparingly soluble: *v. s.* = very soluble; *alm.* = almost; *dec.* = decomposed.

NAME OF CHEMICAL.	WATER.		ALCOHOL.	
	At 15° C. (59° F.)	At 15° C. grs. pr f℥j.	At 15° C. (59° F.)	At 15° C. grs. pr f℥j.
One part is soluble in:—	Parts.		Parts.	
Acid Arsenious	30–80	15.2–5.7	sp.	...
" Benzoic	500	0.9	3	124.6
" Boracic	25	18.2	15	24.9
" Carbolic	20	22.8	v. s.	...
" Chromic	v. s.	...	dec.	...
" Citric	0.75	608	1	374
" Gallic	100	4.5	4.5	83.1
" Salicylic	450	1.01	2.5	149.6
" Tannic	6	76	0.6	623
" Tartaric	0.7	651	2.5	149.6
Alum	10.5	43.4	ins.	...
" Dry, (Exsiccatum)	20	22.8	ins.	...
Aluminium Hydrate	ins.	...	ins.	...
" Sulphate	1.2	380	alm. ins.	...
Ammonium Benzoate	5	91.1	28	13.3
" Bromide	1.5	300.4	150	2 4
" Carbonate	4	114	dec.	...
" Chloride	3	152	1.37	273
" Iodide	1	456	9	41.5
" Nitrate	0.5	910	20	18 7
" Phosphate	4	114	0.5	748
" Sulphate	1.3	350.7	sp.	...
" Valerianate	v. s.	...	v. s.	...
Antimony and Potass. Tartrate	17	26.8	ins.	...
" Oxide	alm. ins.	...	ins.	...
" Sulphide	ins.	...	ins.	...
Apomorphine Hydrochloride	6.8	67	50	7.4
Silver Cyanide	ins.	...	ins.	...
" Iodide	ins.	...	ins.	...
" Nitrate	0.8	570	26	14.3
" Nitrate (fused)	0.6	760	25	14.9
" Oxide	v. sp.	...	ins.	...
Arsenic Iodide	3.5	130.2	10	37.4
Atropine	600	0.7	v. s.	
" Sulphate	0.4	1152	6.5	57.5

SOLUBILITY OF THE MOST IMPORTANT CHEMICALS USED IN MEDICINE, IN WATER AND ALCOHOL. (*Continued.*)

NAME OF CHEMICAL.	WATER.		ALCOHOL.	
	At 15° C. (59° F.)	At 15° C. grs. pr. f℥j.	At 15° C. (59° F.)	At 15° C. grs. pr. f℥j.
Bismuth Citrate	ins.	...	ins.	...
“ and Ammon. Citrate	v. s.	...	sp.	...
“ Subcarbonate	ins.	...	ins.	...
“ Subnitrate	ins.	...	ins.	...
Bromine	33	13.8	dec.	...
Caffeine	75	6.08	35	10.6
Calcium Bromide	0.7	651	1	374
“ Carbonate	ins.	...	ins.	...
“ Chloride	1.5	300.4	8	46.7
“ Hypophosphite	6.8	67	ins.	...
“ Phosphate (precipitated)	ins.	...	ins.	...
Lime (Calx)	750	0.6	ins.	...
Camphor, Monobromated	alm. ins.	...	v. s.	...
Cerium Oxalate	ins.	...	ins.	...
Chloral	v s.	...	v. s.	...
Cinchonidine Sulphate	100	4.5	71	5.2
Cinchonine	alm. ins.	...	110	3.4
“ Sulphate	70	6.5	6	62.3
Codeine	80	5.7	v. s.	...
Chalk (Creta)	ins.	...	ins.	...
Copper Acetate	15	30.2	135	2.7
“ Sulphate	2.6	173.1	ins.	...
Elaterium	ins.	...	125	2.9
Ferric Chloride	v. s.	...	v. s.	...
“ Citrate	s.	...	ins.	...
“ and Ammon. Citrate	v. s.	...	ins.	...
“ “ “ Sulphate	3	152	ins.	...
“ “ “ Tartrate	v. s.	...	ins.	...
“ “ Potass. Tartrate	v. s.	...	ins.	...
“ “ Quinine Citrate	s.	...	ins.	...
“ “ Strychnine Citrate	v. s.	...	ins.	...
“ Hypophosphite	sp.	...	ins.	...
“ Lactate	40	11.4	alm. ins.	...
“ Oxalate	sp.	...	ins.	...
“ Hydrate	ins.	...	ins.	...
“ Phosphate	v. s.	...	ins.	...
“ Pyrophosphate	v. s.	...	ins.	...
“ Sulphate	1.8	253.3	ins.	...
“ Valerianate	ins.	...	v. s.	...
Hyoscyamine Sulphate	v. s.	...	v. s.	...
Iodoform	ins.	...	80	4.6

SOLUBILITY OF THE MOST IMPORTANT CHEMICALS [illegible] IN [illegible] WATER AND ALCOHOL.—[illegible]

NAME OF CHEMICAL	WATER.		ALCOHOL.	
	At 15° C. (59° F.)	[illegible]	[illegible]	[illegible]
Iodine	sp.	--	[illegible]	[illegible]
Lithium Benzoate	4	[illegible]	[illegible]	[illegible]
" Bromide	v. s.	--	v. s.	--
" Carbonate	180	[illegible]	ins.	--
" Citrate	5.5	[illegible]	[illegible]	--
" Salicylate	v. s.	--	v. s.	--
Magnesium Oxide	alm. ins.	--	ins.	--
" Carbonate	alm. ins.	--	ins.	--
" Sulphate	0.8	570	ins.	--
" Sulphite	20	22.5	ins.	--
Manganesium Dioxide (Black Oxide).	ins.	--	ins.	--
Manganesium Sulphate	0.7	651	ins.	--
Mercuric Chloride	16	285	3	124
Mercurous Chloride	ins.	...	ins.	...
Mercuric Cyanide	12.8	8.5	15	24.9
" Iodide (Red)	alm. ins.	...	130	2.8
Mercurous Iodide (Green)	alm. ins.	...	ins.	...
Mercuric Oxide	ins.	...	ins.	...
" Subsulphate	ins.	...	ins.	...
" Sulphide (Red)	ins.	...	ins.	...
Morphine	v. sp.	...	100	3.7
" Acetate	12	38	68	5.5
" Hydrochloride	24	19	63	5.9
" Sulphate	24	19	702	0.5
Phosphorus	ins.	...	v. sp.	...
Physostigmine Salicylate	130	8.5	12	31.1
Picrotoxine	150	8.4	10	37.4
Pilocarpine Hydrochloride	v. s.	...	v. s.	...
Piperine	alm. ins.	...	30	12.4
Plumbic Acetate	1.8	253.8	8	46.7
" Carbonate	ins.	...	ins.	...
" Iodide	2000	0.2	v. sp.	...
" Nitrate	2	228	alm. ins.	...
" Oxide	ins.	...	ins.	...
Potassium Hydrate	0.5	910	2	187
" Acetate	0.4	112.5	2.5	149.6
" Bicarbonate	3.2	142.5	alm. ins.	...
" Bichromate	10	45.6	ins.	...
" Bitartrate	210	2.1	v. sp.	...
" Bromide	1.6	285	200	1.8

SOLUBILITY OF THE MOST IMPORTANT CHEMICALS USED IN MEDICINE, IN WATER AND ALCOHOL. (*Continued.*)

NAME OF CHEMICAL.	WATER.		ALCOHOL.	
	At 15° C. (59° F.)	At 15° C. grs. pr f℥j.	At 15° C. (59° F.)	At 15° C. grs. pr f℥j.
Potassium Carbonate	1	456	ins.	...
" Chlorate	16.5	27.6	v. sp.	...
" Citrate	0.6	760	v. sp.	...
" Cyanide	2	228	sp.	...
" and Sodium Tartrate	2.5	182.4	alm. ins.	...
" Ferrocyanide	4	114	ins.	...
" Hypophosphite	0.6	760	7.3	51
" Iodide	0.8	570	18	20.7
" Nitrate	4	114	alm. ins.	...
" Permanganate	20	22.8	dec.	...
" Sulphate	9	50.6	ins.	...
" Sulphite	4	114	sp.	...
" Tartrate	0.7	651	alm. ins.	...
Quinidine Sulphate	100	4.5	8	46.7
Quinine	1600	0.2	6	62.3
" Bisulphate	10	45.6	32	11.6
" Hydrobromate	16	28.5	3	124.6
" Hydrochloride	34	13.4	3	124.6
" Sulphate	740	0.6	65	5.7
" Valerianate	100	4.5	5	74.8
Sugar, Cane	0.5	910	175	2.1
" Milk	7	65.1	ins.	...
Salicin	28	16.2	30	12.4
Santonin	alm. ins.	...	40	9.3
Sodium Hydrate	1.7	268.2	v. s.	...
" Acetate	3	152	30	12.4
" Arseniate	4	114	v. sp.	...
" Benzoate	1.8	253.3	45	8.3
" Bicarbonate	12	38	ins.	...
" Bisulphite	4	114	72	5.2
" Borate (Borax)	16	28.5	ins.	...
" Bromide	1.2	380	13	28.7
" Carbonate	1.6	285	ins.	...
" Chlorate	1.1	44.5	40	9.3
" Chloride	2.8	162.8	alm. ins.	...
" Hypophosphite	1	456	30	12.4
" Hyposulphite	1.5	300.4	ins.	...
" Iodide	0.6	.760	1.8	207
" Nitrate	1.3	350.7	sp.	...
" Phosphate	6	76	ins.	...
" Pyrophosphate	12	38	ins.	...

SOLUBILITY OF THE MOST IMPORTANT CHEMICALS USED IN MEDICINE, IN WATER AND ALCOHOL. (*Continued.*)

NAME OF CHEMICAL.	WATER.		ALCOHOL.	
	At 15° C. (59° F.)	At 15° C. grs. pr f℥j.	At 15° C. (59° F.)	At 15° C. grs. pr f℥j.
Sodium Salicylate	1.5	300.4	6	62.3
" Santoninate	3	152	12	31.1
" Sulphate	2.8	162.8	ins.	...
" Sulphite	4	114	sp.	...
" Sulphocarbolate	5	91.2	132	2.8
Strychnine	6700	0.06	110	3.4
" Sulphate	10	45.6	60	6.2
Sulphur	ins.	...	ins.	...
Thymol	1200	0.3	1	374
Veratrine	v. sp.	...	3	124.6
Zinc Acetate	3	152	30	12.4
" Bromide	v. s.	...	v. s.	...
" Carbonate	ins.	...	ins.	...
" Chloride	v. s.	...	v. s.	...
" Iodide	v. s.	...	v. s.	...
" Oxide	ins.	...	ins.	...
" Phosphide	ins.	...	ins.	...
" Sulphate	0.6	760	ins.	...
" Valerianate	100	4.5	40	9.3

INCOMPATIBLES.

Substances are said to be chemically incompatible when, on being mixed together, they react upon each other so as to cause an entire change in the properties of the substances so mixed. They may cause the evolution of a gas, an explosive mixture or compound, a poisonous or very active substance formed from comparatively inert ones, or, a precipitation of one or the other of the ingredients in the new compounds formed.

Sometimes two or more substances are brought together with the intent of producing a new substance different from either; as, $2KI + HgCl_2 = HgI_2 + 2KCl$. This can hardly be regarded as an incompatible mixture.

Of physiological or therapeutical incompatibility we shall have nothing to say here. The student will find the following rules of value to him in the beginning.

1. A free acid is incompatible with the alkaloids and the metallic hydrates and carbonates. The three mineral acids displace the organic acids from their salts. The converse of these statements is also true; *i. e.*, metallic hydrates and carbonates are incompatible with the acids.

2. If two substances, when mixed, can form an insoluble third body, or can react so as to generate a gas, they are incompatible.

A knowledge of the solubility of the ordinary salts is, therefore, of great importance to the physician. For example, lead or barium cannot exist in a solution with a sulphate; silver, lead, or mercurous mercury cannot exist in a solution with a chloride.

Substances are, therefore, incompatible with their tests and antidotes.

3. The alkaline hydrates and carbonates are incompatible with the salts of the alkaloids and most salts of the heavy metals.

4. Iodides and bromides precipitate most of the heavy metals, and are therefore incompatible with them.

5. The vegetable astringents and bitters owe these properties largely to their gallic and tannic acids. Tannic acid and most vegetable astringents precipitate the heavy metals from their salts, and are therefore incompatible with them.

6. Powerful oxidizing agents (strong nitric acid, potassium permanganate, hydrogen peroxide, chlorine, the hypochlorites, potassium chlorate, etc.), should not be mixed with easily oxidizable organic substances for fear of forming explosive compounds.

7. The two principal solvents of the U. S. P. are alcohol and water. Each of these has its own class of easily soluble bodies. These bodies are often precipitated from their solutions in either of these solvents by the addition of the other. Thus the tinctures of iodine, camphor, essential oils, the gums and gum-resins, aloes, etc., are precipitated or rendered unsightly by the addition of water or watery solutions of drugs or chemicals.

8. There are some solutions which should always be prescribed alone, or in a plain watery solution, as they readily decompose. Among these may be mentioned the compound syrup of the Hypophosphites, Fowler's, Donovan's, and Lugol's solutions.

These few rules will serve to call the student's attention to the subject, and to the general range of incompatibles.

GLOSSARY

OF UNUSUAL CHEMICAL TERMS.

[The figures in parentheses refer to the pages of this book where a fuller explanation may be found.]

ACTINISM. The chemical effects of light.

Ærugo. Verdigris. Impure subacetate of copper.

Æthiops. Black sulphide of mercury. Hg_2S.

Aerometer. Hydrometer. (13.)

Alabaster. A light-colored, compact gypsum. $CaSO_4$. (197.)

Alchemy. The Arabic name for chemistry, which formerly arose out of the search for the philosopher's stone and elixir of life.

Alembic. A form of still or retort, used in sublimation.

Alkarsin. Oxide of cakodyl, or cacodylic acid. $As(CH_3)_2O_2H$.

Alloy. A mixture or compound formed by fusing two or more metals together.

Amidon. Starch. (271.)

Amorphous. Without a definite crystalline form.

Anhydride. An oxide which can combine with the elements of water to produce an acid. Hence, an acid deprived of one or more molecules of water.

Anode. The + pole of a voltaic circuit.

Apple Oil. Valerianate of amyl.

Aqua Fontana. Aqua, U. S. P.

Aqua Fortis. Crude nitric acid. (141.)

Aqua Phagedænica. Yellow wash. Mercuric hydrate.

Aqua Regia. Nitro-muriatic acid. (142.)

Aqua Vitæ. Brandy.

Argols. Crude cream of Tartar. (193.)

Auripigmentum. Orpiment. Arsenious sulphide.

Arrack. A spirituous drink made from the juice of the cocoanut tree.

Austral. The south pole of a magnet.

Azote. (Fr.) Nitrogen.

Azotic Acid. Nitric acid. (141.)

BALDWIN'S Phosphorus. Fused calcium nitrate, possibly luminous calcium sulphide.

Balsam of Sulphur. A solution of S. in Ol. olivæ.

Barilla. The ashes of sea plants, and of **Salsola Soda.**

Basyl. A term applied to an electro-positive radical.

Battery. An apparatus for the production of electricity by chemical action. (40.)

Baumé. The name of the inventor of a hydrometer bearing this name.

Bell Metal. An alloy of 6 parts copper and 2 parts tin.

Bestuchuf's Tincture. An ethereal solution of Fe_2Cl_6.

Bibron's Antidote. A solution of $HgCl_2$,KI, bromine, alcohol and water.

Bittern. The mother liquor remaining after extracting NaCl from sea water, by evaporation and crystallization.

Black Ash. Impure Na_2CO_3, mixed with carbon.

Black Flux. Made by igniting cream of tartar with one-half its weight of nitre; KNO_3. It contains carbon and K_2CO_3.

Black Drop. Acetum opii. Vinegar of opium.

Black Lead. Plumbago, a native variety of carbon, used for making lead pencils, crucibles, and stove polish. (166.)

Black Salts. The ley of wood-ashes evaporated nearly to dryness.

Black Wash. Contains suboxide of mercury, Hg_2O. (231.)

Bleaching Powder. Chloride of lime. A mixture of chloride and hypochlorite of calcium. (199.)

Blende. Native sulphide of zinc. ZnS. (205.)

Blue Mass. Pilulæ Hydrargyri.

Blue Ointment. Unguentum Hydrargyri.

Blue Vitriol, or Bluestone. Sulphate of copper. (227.)

Bole. An argillaceous earth.

Bone Ash. Bone black, or charred bones. (200.)

Borax. Biborate of sodium. Tetraborate of sodium, $Na_2B_4O_7$. (185.)

Boreal. The north pole of a magnetic needle.

Brass. An alloy of copper and zinc.

Brimstone. Roll sulphur. (120.)

British Barilla. Black Ash.

British Gum. Dextrin. (272.)

Bronze. An alloy of copper and tin. (239.)

Brunswick Green. Oxychloride of copper.

Bunsen Burner. A gas burner used for production of heat. It mixes the gas and air before burning them.

Burnett's Disinfecting Fluid. Solution of $ZnCl_2$. (133.)

Butter of Zinc, Antimony, and Bismuth. Their chlorides.

CALAMINE. Impure, native carbonate of zinc. (205.)

Calcareous Spar. Calcite. $CaCO_3$. (200.)

Calcedony. A native form of SiO_2.

Calcining. Igniting a substance in the air, so as to burn off any oxidizable material, or expel volatile products.

Calcined Mercury. Mercuric oxide. HgO.

Calomel. Mercurous chloride. Mild chloride of mercury. Hg_2Cl_2. (230.)

Caloric. Old term for heat.

Calorie. The unit of heat used in determining the heat of combination of chemical compounds. (27.)

Camphene. Oil of turpentine. Camphene burning fluid is a solution of turpentine in alcohol.

Canton's Phosphorus. Luminous CaS, or luminous paint.

Caput Mortuum. The residue left after ignition of $FeSO_4$ or iron pyrites. Impure Fe_2O_3. (215.)

Caramel. Burned sugar. (269.)

Carbolic Acid. Phenic acid. Phenyl alcohol. (287.)

Carburet. Carbide.

Catalysis. The action of a body in promoting combination or decomposition by its presence, the body itself remaining unchanged.

Cathode. The negative pole of a galvanic circuit.

Chalk. An amorphous carbonate of lime. (200.)

Chameleon Mineral. Permanganate of potassium.

Choke Damp. Carbonic anhydride. CO_2. (171.)

Chrome Green. A mixture of chrome yellow and Prussian blue; or sesquioxide of chromium. Cr_2O_3. (208.)

Chrome Vermilion. Dichromate of lead. $PbCr_2O_7$.

Chrome Yellow. Chromate of lead. $PbCrO_4$. (223.)

Cinnabar. Native red sulphide of mercury. HgS. (229.)

Citrine Ointment. Nitrate of mercury ointment.

Clay. Impure silicate of alumina.

Clay Ironstone. A variety of hematite iron ore. Fe_2O_3.

Colcothar. Ferric oxide. Fe_2O_3. Rouge, crocus.

Collodion. Solution of gun cotton in alcohol and ether. (273.)

Colloids. Jelly-like or non-crystallizable bodies. (57.)

Colophony. Common resin, or rosin. (252.)

Condy's Solution. Contains permanganate of potassium. $K_2Mn_2O_8$. (133.)

Common Salt. Sodium chloride. $NaCl$.

Copperas. Green vitriol. Crystallized ferrous sulphate. $FeSO_4.7Aq$. (214.)

Corrosive Sublimate. Mercuric chloride. Bichloride of mercury. (326.)

Cream of Tartar. Acid potassium tartrate. $HKC_4H_4O_6$. (193.)

Crocus of Antimony, or Crocus Metallorum. Oxysulphide of antimony. (159.)

Crocus Martis. Colcothar. Fe_2O_3.

Crystals of Venus. Copper acetate. $Cu(C_2H_3O_2)_2.H_2O$.

Crystalloids. Crystallizable bodies, as distinguished from colloids. (58.)

Cubic Nitre. Sodium nitrate. $NaNO_3$.

Cupellation. The process of purifying silver or gold in a **cupel** or cup made of bone ash. When the alloy is strongly heated in air, the other metals oxidize and the cupel absorbs the oxide, leaving the silver or gold pure.

DECANTATION. The process of pouring off the clear liquid above a sediment.

Decoction. An extract of an organic substance, made with boiling water.

Displacement. Expelling a fluid from a vessel by another of different density.

Decrepitation. The crackling of certain salts when suddenly heated.

Deflagration. A rapid and scintillating combustion. It takes place in certain mixtures containing the nitrates or chlorates.

Deliquescent. An adjective applied to those substances which attract moisture from the air and liquefy. (55, 59.)

Destructive Distillation. Dry distillation, conducted with the object of destroying the substance and producing new ones. (25.)

Detonation. Rapid chemical action, accompanied by flame and noise. An explosion.

De Valangin's Arsenical Solution. A solution of the chloride. $AsCl_3$. **Liq. Arsen. Hydrochlor.**

Dew-point. The temperature at which the moisture of the air begins to deposit.

Dialysis. The process of the diffusion of liquids and solutions through membranes. (57.)

Dimorphous. Crystallizing in two distinct systems. (58.)

Donovan's Solution. Contains the iodides of arsenic and mercury.

Dolomite. Magnesium limestone. (203.)

Dover's Powder. Compound ipecac. powder; contains opium. (1 gr. in 10).

Drummond Light. Calcium light.

Dry Distillation. The process of subjecting solid or organic bodies to heat, in a closed retort.

Ductile. Capable of being drawn into wire, or rolled out into sheets.

Dutch Gold. A species of brass, usually sold in very thin leaves or sheets.

Dutch Liquid. Ethene dichloride. $C_2H_4Cl_2$.

Dutch White. Impure white lead.

EAU de Javelle. A solution of chlorinated potassa, or potass. hypochlorite. KOCl.

Educts. The proximate principles of which bodies are supposed to be formed.

Effervescence. The rapid escape of gas from a liquid.

Efflorescence. The escape of water of crystallization and the consequent crumbling down of the crystal. (59.)

Electrode. The pole or wire forming a part of a voltaic circuit. (45.)

Electrolysis. Decomposition by means of a strong electric current. (51.)

Element. A substance which has never been decomposed.

Elixir of Vitriol. Aromatic sulphuric acid.

Elutriation. The process of separating the finer and lighter particles of a powder from the coarser, by suspending them in water and pouring off the lighter floating particles with the water.

Emerald Green. Schweinfurth green, or aceto-arsenite of copper. See Paris green.

Emery. An impure corundum. Al_2O_3. (238.)

Eosine. Tetrabromo-fluorescine. A beautiful, red, artificial coloring matter.

Epsom Salt. $MgSO_4.7Aq$. (204.)

Eremacausis. The slow decay of organic substances in the air.

Essence of Mirbane. Nitro-benzol. (293.)

Essential Oils. Volatile oils. (277.)

Eudiometer. A graduated glass tube, closed at one end, used for measuring gases.

FERMENTATION. (321.)

Filter. A porous substance used to separate a solid and liquid by allowing the latter to pass through, while the former is retained.

Fire Damp. Light-carbureted hydrogen (Marsh gas), mixed with air. (247.)

Fixed Air. Choke damp. CO_2. (171.)

Flint. An impure variety of silica. SiO_2.

Flowers of Antimony. Oxide of antimony.

Flowers of Benzoin. Benzoic acid. (289.)

Flowers of Sulphur. Sulphur sublimatum. U. S. P. (120.)

Flowers of Zinc. Oxide of zinc. ZnO. (206.)

Fluorescence. The property possessed by certain bodies, as quinine salts, of rendering visible the ultra-violet or chemical rays of the solar spectrum.

Fluor Spar. Native calcium fluoride. CaF_2. (111.)

Flux. A material added to ores in smelting, to form an easily fusible slag.

Foliated Earth of Tartar. $KC_2H_3O_2$. Potassium acetate.

Fool's Gold. Iron pyrites.

Fowler's Solution. Solution of the arsenite of potassium. (152.)

French Chalk. Silicate of magnesia. Soapstone; talc.

Fructose. Fruit sugar.

Fuchsine. Aniline red. Magenta. $C_{20}H_{19}N_3$.

Fuming Liquor of Libarius. Solution of stannic chloride.

Fusel Oil. Amylic alcohol. (264.)

Fusible Calculus. One composed of a mixture of phosphate of lime, and ammonio-magnesium phosphate.

Fusible Metal. Bismuth 2 parts, lead 1 part, and tin 1 part. Melts at about 200°F.

GALENA. Native lead sulphide. (219.)

Galvano-Cautery. A surgical knife heated by a galvanic current.

German Silver. An alloy of copper, nickel and zinc.

Glass. An artificial silicate of calcium, sodium, iron, lead, etc.

Glass of Antimony. Fused trisulphide of antimony. Sb_2S_3.

Glass of Borax. Fused borax.

Glauber's Salt. Sodium sulphate. (182.)

Glucose. Grape sugar. Now made on large scale, from corn starch. (267.)

Glucoside. (See p. 297.)

Glyceroles and Glycerita. Simple glycerine solutions.

Goniometer. An instrument for measuring the angles of crystals.

Goulard's Extract, and Cerate. Contain subacetate of lead.

Graphine. Carbon deposited in gas retorts.

Graphite. Plumbago. Black lead. (166.)

Green Vitriol. Copperas. $FeSO_4.7Aq$. (214.)

Guano. A deposit of excrement of sea fowl.

Gypsum. Calcium sulphate. Plaster of Paris. (199.)

HALOGEN Elements. Haloid Salt. The elements of the chlorine group and their binary compounds. (111.)

Harle's Solution. Solution of arsenite of sodium.

Haschisch. Indian hemp.

Hartshorn. Ammonia. (134.)

Hepar Sulphuris. Liver of sulphur. Potassium sulphide.

Hive Syrup. Compound syrup of squills.

Hoffmann's Anodyne. Spirit. Æther. Co. Ether, 1 pint; alcohol, 1 pint; ethereal oil, f℥ vj.

Homberg's Pyrophyrus. Is made by igniting potassium, alum and charcoal.

Homologues. (245.)

Homologous Series. (245.)

Huxham's Tincture. Compound tincture of cinchona.

Hydracid. A binary acid. Contains no oxygen.

Hydrate. A compound containing hydroxyl, HO, combined to a positive radical. (77.)

Hydrochloride. A compound of HCl, formed by union of the whole molecule by synthesis; as compounds with the alkaloids.

Hygrometer. An instrument for the determination of the relative amount of moisture in the air.

ICE Vinegar. Glacial acetic acid. (279.)

Incandescence. The glow of a highly heated body.

Incineration. The reduction of a substance to ashes by burning.

Incompatible. Incapable of being mixed without chemical change. (343.)

Infusion. An extract of an organic substance, made by pouring hot water upon it and allowing it to stand for some hours.

Inosite. Muscle sugar. (271.)

Ion. A body going to the positive (anode) or negative (cathode) pole of a galvanic battery during electrolysis.

Iron Pyrites. Native sulphide of iron. Fools' gold.

Isinglass. A variety of gelatin, or fish glue. Sometimes erroneously applied to mica.

Ivory Black. Animal charcoal, made by distilling ivory scraps; is now generally applied to bone black.

JAMES' Powder. Antimonial powder.

Japan Black. A varnish composed of asphaltum, turpentine, linseed oil and umber.

Jesuits' Powder. Powdered cinchona bark.

KAOLIN. A pure white clay.

Kelp. Ashes of sea weeds; used as a source of iodine and carbonate of sodium.

Kermes' Mineral. Sb_2S_3. (160.)

King's Yellow. Orpiment. As_2S_3.

Kyan's Disinfectant. Solution of $HgCl_2$.

LABARRAQUE'S Disinfecting Liquid. Solution of hypochlorite of sodium or chlorinated soda. (186.)

Lac Sulphuris. Precipitated sulphur. (122.)

Lactin—Lactose. Sugar of milk. (269.)

Lacquer. A varnish used for brass, etc.

Lady Webster Pill. Pill of aloes.

Lake. An organic coloring matter precipitated with aluminium hydrate. Used as pigments.

Lampblack. The soot of burning turpentine. (167.)

Lapis Infernalis. Lunar caustic. $AgNO_3$.

Laughing Gas. Nitrous oxide. N_2O. Dentists' gas. (137.)

Lead Water. Diluted Goulard's Extract, containing subacetate of lead. (223.)

Ledoyen's Disinfecting Liquid. Solution of nitrate of lead.

Levigation. The reduction of a substance to an impalpable powder, by rubbing on a slab, with sufficient water to form a paste, with a flat pestle called a muller.

Lime. CaO. (197.)

Limestone. A native carbonate of lime. (197.)

Litharge. PbO. Semi-vitrified oxide of lead. (220.)

Lithic Acid. Uric Acid. (291.)

Liver of Sulphur. Potass. sulphuret. K_2S. (192.)

Lixiviation. The separation of the soluble portions of a substance by causing water to filter through it. (110.)

Loadstone. The native magnetic oxide of iron, or magnetite. A magnet. (205.)

Lugol's Solution. Compound solution of iodine. Iodine held in solution by KI. (118.)

Lunar Caustic. Nitrate of silver. $AgNO_3$. (236.)

Lute. An adhesive mixture for closing the joints of apparatus, to prevent the escape of vapors, etc.

MACERATION. The long continued soaking of a substance in water at common temperatures. (110.)

Macquer's Salt. Potassium arseniate.

Magendie's Solution. Morph. sulphate, gr. xvj, water, f℥j.

Magistery of Bismuth. Subnitrate.

Magma. A pasty mass.

Magnesia Alba. Magnesium carbonate. (205.)

Malleable. Capable of being worked under the hammer.

Marble. Nearly pure native carbonate of lime. (197.)

Marine Acid. Muriatic acid. HCl. (114.)

Martial Æthops. Fe_3O_4. Magnetic oxide of iron.

Massicot. Amorphous oxide of lead. PbO. Powdered litharge. (220.)

Matrass. A glass vessel with a long neck, or a tube sealed at one end. Used for heating dry substances.

Mercaptan. An alcohol in which O is replaced by sulphur.

Metalloid. Non-metal. (89.)

Menstruum. A solvent, or medium of chemical reaction.

Metameric Bodies. The same as isomeric. (246.)

Microcosmic Salt. $NaNH_4HPO_4$.

Milk of Sulphur. Precipitated sulphur.

Milk of Lime. Whitewash. (198.)

Mineral Water. Water charged with carbonic acid; also natural water holding medicinal substances in solution. (109.)

Mineral Yellow. Oxychloride of lead.

Minium. Red oxide of lead. $2PbO.PbO_2$. (221.)

Molecule. (See page 63.)

Monsel's Salt. Subsulphate of iron. (216.)

Mosaic Gold. Brass.

Mordant. A substance used to fix colors on fabrics.

Mountain Blue. Azurite. Native basic carbonate of copper.

Mountain Green. Malachite. A native basic carbonate of copper. (228.)
Mulberry Calculus. Calcium oxalate. CaC_2O_4.
Muriatic Acid. Hydrochloric acid. (114.)
Muriate. A chloride.

NAPHTHA. A light hydrocarbon obtained from petroleum, and boiling at about 80° to 105°C. (248.)
Natron. Native carbonate of sodium. (183.)
Neutral. Without action on test paper. Neither acid nor alkaline.
Neutral Mixture. Solution of citrate of potassium.
Nitre. Saltpetre. KNO_3. (190.)
Normal Salt. Neither acid nor basic.

OBSIDIAN. Volcanic glass.
Ochre. A native mixture of clay and ferric oxide, used as a paint.
Oil of Vitriol. H_2SO_4. Sulphuric acid. (125.)
Oil of Wine. Ethyl sulphate. $(C_2H_5)_2SO_4$.
Oreide. A species of brass resembling gold, and used for jewelry.
Orpiment. Arsenious sulphide. As_2S_3. (151.)
Ortho-acid. An acid in which each bond of the kernel is united to hydroxyl. (OH.)
Osmosis. The diffusion of liquids through porous septa. See Dialysis. (57.)
Ox-acid. A ternary acid containing oxygen.

PACKFONG. A variety of German silver.
Particle. A minute portion of matter.
Paris Green. Impure Schweinfurth green. Aceto-arsenite of copper. (152.)
Pearl Ash. Impure carbonate of potassium. (191.)
Pearl Powder. Subnitrate, or oxychloride of bismuth. ($BiONO_3$ or $BiOCl$.)
Pearl White. $BiONO_3$, or $BiOCl$.
Pearson's Salt. Arseniate of sodium. Na_3AsO_4.
Pewter. An alloy of variable composition. Usually composed of tin, lead, copper and antimony, or zinc.
Pinchbeck Gold. A species of brass.
Plaster of Paris. Calcium sulphate. Calcined gypsum. (199.)

Platinum Black and Sponge. Finely divided platinum. (240.)
Plumbago. Native carbon. Graphite. (166.)
Potash. Impure carbonate of potassium. (191.)
Potassa. Oxide or hydrate of potassium.
Powder of Algaroth. Oxychloride of antimony. (159.)
Precipitate. An insoluble substance, formed on bringing two or more substances together in solution.
Precipitatum Per Se. Mercuric oxide. HgO. Made by heating mercury to near its boiling point until it oxidizes.
Preston Salts. Carbonate of ammonia, flavored with some essential oil.
Prussic Acid. Hydrocyanic Acid. (176.)
Prussian Blue. Ferric ferrocyanide. (179.)
Pseudomorph. A mineral crystallized in the form that belongs to another mineral.
Puce-oxide of Lead. Lead peroxide, or brown oxide.
Purgative Mineral Water. Liq. magnes. citrat.
Purple of Cassius. A pigment produced by treating chloride of gold with a solution of stannous chloride. (238.)
Putty. Composed of whiting and linseed oil.
Putty Powder. Stannic oxide.
Pyrites. Native sulphide of iron.
Pyroxilic Spirit. Wood alcohol. Methyl alcohol. (259.)
Pyroxylin. Gun cotton. Trinitro-cellulose. (273.)

QUANTIVALENCE. Quantity of combining power; applied to atoms. (71.)
Quartz. SiO_2. (165.)
Quevenne's Iron. Ferrum redactum. (213.)
Quicklime. Caustic lime. CaO. (197.)
Quicksilver. Mercury. (229.)

RADICAL. An atom, or group of atoms, forming the basis of a series of compounds. (74.)
Radical Vinegar. Glacial acetic acid.
Realgar. Red sulphide of arsenic. (151.)
Red Precipitate. Red oxide of mercury. HgO.
Red Prussiate of Potash. Ferricyanide of potassium. (179.)
Red Tartar. Argol. (193.)
Regulus of Antimony. Metallic antimony.
Rochelle Salt. Tartrate of potassium and sodium. (193.)

Roche and Roman Alums. Are varieties of potassium alum.

Rock Crystal. Quartz. SiO_2. (165.)

Roman Vitriol. Sulphate of copper. $CuSO_4$.

Rouge. Ferric oxide in fine powder. Fe_2O_3.

Ruby. Native Al_2O_3, of a beautiful red color. (238.)

Rust. Ferric oxide; generally containing some ferric hydrate. (213.)

SACCHARUM Saturni. Acetate of lead. (223.)

Safety Lamp. A lamp inclosed in wire gauze, to prevent explosions in explosive gases in mines, cellars, etc.

Sal Æratus. Potassium bicarbonate. (191.)

Sal Alembroth. Double chloride of mercury and ammonium.

Sal Ammoniac. Ammonium chloride. (135.)

Sal Diureticus. Potassium acetate.

Sal Enixum. Potassium bisulphate.

Sal Mirabile. Sodium sulphate.

Sal Perlatum. Sodium phosphate.

Sal Prunelle. Fused nitre. KNO_3.

Sal Volatile. Ammonium carbonate.

Salt of Lemon and Salt of Sorrell. Potassium binoxalate. (192.)

Salt of Phosphorus. Microcosmic salt.

Salt of Tartar. Pure potassium carbonate. (191.)

Saltpetre. Potassium nitrate. (190.)

Sapphire. A native form of Al_2O_3.

Scheele's Green. Arsenite of copper. (152.)

Schlippe's Salt. Sodium sulphantimoniate. Na_3SbS_4.

Schweinfurth Green. Copper aceto-arsenite. (227.)

Seidlitz Powder. A mixture of sodium bicarbonate and Rochelle salt in one paper and tartaric acid in another.

Seignette's Salt. Rochelle salt. $KNaC_4H_4O_6$. (193.)

Sizing. A gelatinous mixture put into paper or cloth, to fill up the pores.

Slag. The fused impurities from smelting of ores. (217.)

Smalt. Glass colored blue by oxide of cobalt and powdered.

Smelting. The process of recovering the metals from their ores.

Soda Ash. Crude sodium carbonate. (184.)

Soda Saltpetre. Sodium nitrate. $NaNO_3$. (185.)

Soluble Tartar. Neutral potassium tartrate. (192.)

Speculum Metal. An alloy of copper and tin.

Speiss. Impure, fused nickel arsenide.
Spelter. Commercial zinc.
Spermaceti. A fat obtained from the sperm whale.
Spirit of Hartshorn. Spirit of ammonia. Solution of ammonia in alcohol.
Spirit of Mindererus. Solution of ammonium acetate. (196.)
Spirit of Nitre. Spirit of nitrous ether. (141.)
Spirit of Salt. Muriatic acid. (114.)
Soapstone. Talc. (203.)
Soda Water. Water artificially charged with CO_2 under pressure.
Solder. An alloy of tin and lead.
Soluble Glass. See Water Glass. (165.)
Spirit of Wine. Alcohol.
Steinbuhl Yellow. Barium chromate. $BaCrO_4$.
Substitution. The displacement of an atom in a molecule by another atom of a different kind. (254.)
Sugar of Lead. Lead actetate. (223.)
Sulphuret. Sulphide.
Sulphuric Æther. Ethylic ether. $C_4H_{10}O$. (275.)
Sulphur Vinum. Impure sulphur. Horse brimstone.

TALMI Gold. An alloy of copper and aluminium (90 to 10).
Tartar Emetic. Antimony and potassium tartrate. (160.)
Tasteless Purging Salt. Sodium phosphate.
Thénard's Blue. A compound of the oxides of aluminium and cobalt.
Tincal. Native borax. $Na_2B_4O_7$. (185.)
Tincture. A solution in alcohol. When in ether, it is called an ethereal tincture.
Tombac. A kind of brass.
Tournesol. Litmus.
Trituration. Rubbing in a mortar.
Trona. Native sodium carbonate. (183.)
Turnbull's Blue. Ferrous ferricyanide. (179.)
Turner's Cerate. Calamine cerate.
Turner's Yellow. Lead oxychloride.
Turpeth Mineral. Yellow sulphate of mercury. (233.)
Tutty. Impure zinc oxide.
Type Metal. An alloy of lead and antimony. (1 to 3.)

UMBER. A native silicate of aluminium, with oxides of iron and manganese. Used as a brown paint.

Ultramarine. Lapis lazuli. A compound of aluminium sodium silicate with sodium sulphides. A beautiful blue pigment. It is now prepared artificially, as well as a green, red and violet variety.

VALENCE of Atoms. Quantity of combining power.

Vallet's Mass. $FeCO_3$ made into a pill mass.

Varec. Kelp. Ash of sea weeds.

Verd Antique. Precious serpentine.

Verdigris. Impure copper subacetate. (228.)

Verditer. Basic copper carbonate.

Vermilion. Artificial mercuric sulphide. HgS. (233.)

Vitriolic Acid. Sulphuric acid. (125.)

WATER Glass. Soluble glass. Sodium silicate. (166.)

White Arsenic. Arsenious oxide.

White Lead. A basic lead carbonate. (223.)

White Precipitate. Ammoniated mercury. Mercur-amidogen chloride. NH_2HgCl. (232.)

White Vitriol. Zinc sulphate. (206.)

Whiting. Prepared chalk. $CaCO_3$. White clay, often sold for whiting.

Wood Naphtha and Wood Spirit. Methyl alcohol. (278.)

Wood Vinegar. Pyroligneous acid. Impure acetic acid. (278.)

YELLOW Prussiate of Potash. Potassium ferrocyanide. (179.)

Yellow Wash. Made by adding corrosive sublimate to lime water. It forms mercuric oxide. (233.)

ZAFFRE. Impure cobalt oxide.

Zinc White. Zinc oxide. Used as a paint. (223.)

Zymosis. The peculiar action caused by a ferment.

INDEX.

ABSORPTION spectra, 35
Acetamide, 292
Acetone, 258
Acetylenes, 247 249
Acid, acetic, 278
" acetic, glacial, 279
" aconitic, 286
" amido-caproic, 317
" antimonic, 160
" antimonious, 160
" arsenic, 153
" arsenious, 152
" benzoic, 289
" bismuthic, 162
" boracic, 164
" bromic, 117
" butyric, 280
" caffeetannic, 303
" capric, 305
" caproic, 259, 305
" caprylic, 259, 305
" carbazotic, 288
" carbolic, 287
" carbonic, 174
" catechutannic, 303
" chloric, 116
" chlorous, 116
" cholesteric, 313
" cholic, 312
" chromic, 209
" citric, 286
" cyanic, 178
" dextrotartaric, 285
" dichloracetic, 279
" ethylsulphuric, 275
" formic, 278
" gallic, 303
" gallotannic, 303
" glycero-phosphoric, 266
" glycero-sulphuric, 266
" glycollic, 282
" glycocholic, 312
" hippuric, 291
" hydriodic, 118
Acid, hydrobromic, 117
" hydrochloric, 114
" hydrocyanic, 176
" hydrofluoric, 111
" hydrofluo-silicic, 165
" hydrosulphuric, 122
" hyoglycocholic, 312
" hypobromous, 117
" hypochlorous, 115
" hypophosphorous, 148
" hyponitrous, 138
" hyposulphurous, 124
" indigotin-sulphonic, 315
" iodic, 119
" isobutyric, 281
" kinotannic, 303
" kyanuric, 312
" lactic, 259, 282
" lævotartaric, 285
" lithofellic, 312
" lithuric, 312
" malic, 284
" mannitic, 267
" mesotartaric, 285
" metaboric, 165
" metaphosphoric, 148
" metatartaric, 285
" muriatic, 114
" nitric, 141
" " in air, 131
" " fuming, 142
" nitro-hydrochloric, 142
" nitrous, 139
" " in air, 131
" oleic, 282
" orthosulphuric, 318
" osmic, 242
" oxalic, 283
" oxybenzoic, 290
" palmitic, 282
" perbromic, 117
" perchloric, 116
" periodic, 119
" phenic, 287

Acid, phenylsulphuric, 288
“ phosphoric, 147
“ “ glacial, 148
“ phosphorous, 149
“ picric, 288
“ prussic, 176
“ pyroantimonic, 160
“ pyrophosphoric, 148
“ pyrotartaric, 285
“ quercotannic, 303
“ quinic. 303
“ racemic, 285
“ saccharic, 267
“ salicylic, 290
“ silicic, 166
“ stearic, 282
“ succinic, 283
“ sulphocarbolic, 288
“ sulphocarbonic, 175
“ sulphuric, 125
“ “ fuming, 127
“ “ Nordhausen, 127
“ sulphurous, 124
“ sulphydric, 122
“ tannic, 303
“ tartaric, 285
“ taurocholic, 312
“ tetraboric, 165
“ trichloracetic, 257
“ trioxybenzoic, 303
“ tungstic, 242
“ uric, 291
“ valerianic, 281
“ valeric. 281
“ albumin, 309
“ salt, 78
Acids, atomicity of, 259
“ basicity of, 77
“ biliary, 312
“ definition, 76, 258
“ fatty, 278
“ organic, 258, 277
“ sulpho, 77
“ thionic, 124
Aconitine, 297
Actinic rays, 36
Adhesion, 13
Actinism, 36
Æthiops mineral, 233
Air, gases in, 131
Albumin, egg, 306
Albumin, serum, 306
“ vegetable, 306
Albumins, derived, 309
Albuminates, acid, 309
“ alkali, 309
Albuminose, 309
Alcohol, absolute, 262
“ amyl, 264
“ butyl, 259
“ ceryl, 259
“ cetyl, 259
“ commercial. 263
“ diatomic, 256
“ ethyl, 261
“ heptyl, 259
“ hexyl, 259
“ melissyl, 259
“ menthyl, 251
“ methyl, 259
“ octyl, 259
“ phenyl, 287
“ propyl, 259
“ physiological action, 262
Alcohols, hexatomic, 266
“ monatomic, 259
“ table of, 259
“ tetratomic, 266
“ triatomic, 256, 265
Aldehyde, ethyl or acetic, 257
Aldehydes, 257
Algaroth, powder of, 159
Alizarine, 253
Alkali albumin, 309
Alkaline earths, metals of, 180
Alkaloids, 294
“ cadaveric, 300
“ glucoside, 296
“ natural, 294
“ opium, 298
“ putrefactive, 300
Allantoin, 318
Allotropism, 113
Alum, burnt, 240
“ ammonium, 240
“ chrome, 240
“ ferric, 240
“ manganese, 240
“ potassium, 240
Alumina, 239
Aluminates, 239
Aluminium, 238

Aluminium, bronze, 239
" chloride, 239
" hydrate, 239
" oxide, 239
" sulphate, 239
Aluminium and ammon. sulphate, 240
Amalgam, 230
Amalgamation, 42
Amber, 252
Amides, 292
" acid, 292
Amines, 293
Ammonia, 134
" action on economy, 196
Ammoniacum, 252
Ammonium, acetate, 196
" benzoate, 196, 289
" bromide, 195
" carbonate, 195
" chloride, 194
" compounds, 194
" cyanide, 194
" hydrate, 194
" iodide, 195
" molybdate, 241
" muriate, 194
" nitrate, 196
" phosphate, 196
" sulphate, 196
" sulphide, 195
" sulphydrate, 131
" sulphocyanate, 178
Ammonias, compound, 291
Ammonio-ferrous, sulphate 214
Ammonio-magnesium phosphate, 204
Ampére's law, 17, 66
Amygdalin, 301-319
Amyl acetate, 276
Amyl alcohol, 264
Amyl nitrate, 276
Amyloses, 271
Amylum, 271
Analysis, toxicological, 157
Anæsthetics, 326, 327
Anglesite, 222
Anhydride, boric, 164
" molybdic, 241
" nitric, 140
" nitrous, 139
Anhydride, silicic, 165
" sulphuric, 125
" sulphurous, 124
Anhydrite, 199
Anilin, 293, 297
" colors, 294.
Annatto, 316
Anthracene, 258
Antiarin, 301
Antimony, 158
Antimony and potass., tartrate of, 160, 194
Antimony, butter of, 159
" chloride, 159
" crocus of, 159
" glass of, 159
" hydride, 158
" oxide, 160
" oxychloride, 159
" protochloride, 159
" pentachloride, 159
" physiological action, 161.
" sulphate, 160
" sulphide, 159
" trioxide, 160
" trichloride, 159
" vermilion, 160
Antipyrine, 297
Antiseptics, 132
Apomorphine, 298
Appendix, 329
Aqua, 109
" ammoniæ, 135, 194
" " fort., 135, 194
" chlorina, 113
" fortis, 147
" regia, 142
Arabin, 274
Arabinose, 274
Arbutin, 301
Argentum, 235
Argol, 193, 285
Arsenic, 149
" disulphide, 151
" fluoride, 151
" iodide, 151
" oxides, 152
" pentasulphide, 151
" poisoning, 153
" tribromide, 151

Burnett's fluid, 138, 326
Butter, 305
  " of antimony, 159
Butterine, 305
Butylen, 249
Butyric acid, 280
  " fermentation, 322

CADMIUM, 203, 207.
  " chloride, 207
  " compounds of, 207
  " hydrate, 207
  " iodide, 207
  " oxide, 207
  " sulphate, 207
  " sulphide, 207
Cæsium, 194
Caffeine, 297
Calabar bean, 327
Calamine, 205
Calcium, 197
  " bromide, 198
  " carbonate, 200
  " chloride, 198
  " hydrate, 198
  " iodide, 198
  " hopochlorite, 198
  " light, 198
  " oxalate, 201
  " oxide, 197
  " salts, 198
  " sulphate, 199
  " sulphocarbolate, 288
  " physiological effects, 201
Calculus, fusible, 205
Callaud's battery, 41
Calomel, 230
Caloric, 27
Calorie, 27
Calx, 197
Camphor, 250
  " dibromide, 251
  " monobromated, 251
Camphors, 250
Cane sugar, 268
Caoutchouc, 250
Caramel, 269
Carat, 238
Carbohydrates, 267
Carbon, 166
  " amorphous, 167
Carbon and hydrogen, 243
  " and oxygen, 170
  " and sulphur, 175
  " dioxide, 171
  " " in air, 130
  " disulphide, 131, 175
  " gas retort, 168
  " monoxide, 170
  " monosulphide, 175
  " nuclei, 244
  " suboxide, 170
Carbonates, 174
Carmine, 316
Carnallite, 188
Carnin, 318
Casein, 309
Caseins, vegetable, 310
Caustic alkalies, 180
  " " poisoning by, 186
  " potash, 188
  " soda, 181
  " lunar, 236
Cell, Bunsen's, 40
  " galvanic, theory of, 38
  " Groves', 40
  " Leclanché, 41
Cellulose, 273
Cement, 198
Cerasin, 274
Cerebrote, 317
Cerussite, 219, 222
Chalcocite, 225
Chalcopyrite, 225
Chalk, 200
  " prepared, 200
Charcoal, 167
  " animal, 167
  " wood, 167
Charles, law of, 19
Chemical affinity, 14, 38
  " elements, definition of, 63
  " elements, table of, 64
  " equations, 85
  " notation, 71
  " physics, 16
  " reactions, 83
  " " rules for, 85
  " symbols, 71
Chemicals, solubility of, 337
Chemism, 13, 38

Chemistry, definition of, 9
  " inorganic, 89
  " organic, 243
  " theoretical, 63
Chloral, 257, 327
  " hydrate, 257
Chloride of lime, 198
Chlorine, 112, 131
  " oxides, 115
  " hydride, 114
Chloroform, 260
  " physiological effects, 261
  " vapors, 327
Chlorophyll, 316
Cholepyrrhin, 313
Cholesterin, 312
Chondrin, 311
Chromates, 209
  " toxicology, 210
Chrome, alum, 209, 240
  " iron, 208
  " yellow, 223
Chromic anhydride, 208
  " chlorides, 209
  " oxide, 208
  " trioxide, 208
Chromium, 208
  " alums, 209
  " oxides, 208
  " sulphates, 209
Chromous chlorides, 209
[illegible], composition of, 330
Cinchonidine, 298
Cinchonine, 297
Cinnabar, [illegible], 288
[illegible]
[illegible] of elements, 89
[illegible]
  anthracite, 168
  bituminous, 168
[illegible]
[illegible]
  [illegible] of, 219
[illegible] 219
[illegible]
[illegible]
[illegible]
[illegible]
[illegible]
[illegible]
Cohesion, 13
Coil, induction, 46
Coke, 168
Colchicine, 298
Colcothar, 215
Colic, painters', 223
Collagen, 311
Collodion, 273
  " flexible, 273
  " styptic, 273
Colloids, 57
Colophony, 252
Coloring matters, vegetable, 315
Combustion, 95
  " moist, 106
Compound body, 63
Compounds, 76, 317
  " scale, 285
Conductors, resistance of, 43
Conglutin, 310
Coniferin, 301
Conine, 298
Convolvulin, 301
Convulsives, 324
Copal, 252
Copperas, 214
Copper, 224
  " acetates, 228
  " arsenite, 227
  " carbonates, 227
  " physiological action, 228
  " pigments, 228
  " pyrites, 225
  " sulphate, 227, 325
  " stone, 225
Corrosive sublimate, 326
Corundum, 238
Cotton, gun, 273
Cream of tartar, 193
Crocin, 316
Crocoisite, 219
Crystallin, 307
Crystallizations, 57
  " water of, 59
Crystals, system of, 59
  " forms of, 57, 59
Crystalloids, 58
Crystallography, 57
Culinary paradox, 24
Cupric compounds, 226
Cuprite, 225

Cuprous compounds, 225
Currents, extra, 47
" Faradic, 47
" induced, 45–47
" interrupted, 47
" secondary, 47
" thermo-electric, 49
Cyanides, 176
" compound, 179
" toxicology, 177
Cyanogen, 176
Cynogene, 248

DAPHNIN, 296
Decimal system, 14
Definite proportion, law of, 92
Deliquescence, 55–59
Deliriants, 324
Density, 11
Deodorizers, 132
Deoxidation, 98
Deposit, lateritious, 291
Depressants, 324
Dew point, 130
Dextrin, 272
Dextrogyrate, 268
Dextrose, 270
Dialysates, 57
Dialyser, 57
Dialysis, 57
Diamond, 166
Diastase, 318
" pancreatic, 318
Dicalcium phosphate, 200
Diffusion, 21, 55
" of gases, 20
" of liquids, 56
" Graham's law of, 21
Digestion, 110
Digitalin, 302
Digitalis, 327
Dihydro-phenol, 288
Dimethyl oxichinine, 297
Dimorphism, 59
Dipotassium sulphate, 192
Disinfectant, Lanande's, 133
Disinfectants, 132
" commercial, 134
Disodium phosphate, 184
" tetraborate, 185
Distillation, 25
Distillation, destructive, 25
" fractional, 25
Dolomite, 203, 205
Drummond light, 193
Dualine, 266
Dulcite, 266, 267
Dulcose, 267
Dynamite, 266

EBONITE, 250
Efflorescence, 59
Electrical induction, 37
" tension, 40
" current or circuit, 39
Electricity, 36
" effects of, 51
" extra current, 47
" frictional or static, 50
" magneto-, 48
" resistance of conductors of, 43
" secondary or induced currents, of, 45–47
Electro-negative, 89
Electro-metallurgy or plating, 52
Electro-motor force, 40
Electro-positive, 89
Electrodes, 45
Electrolysis, 51
Elements, chemical, definition, 63
" classification of, 89
" non-metallic, 90
" typical, 91
Elutriation, 200
Emery, 238
Emetics, 325
Emulsin, 318, 319
" ferment, 318
Enzymes, 318
Epsom salt, 204
Equations, 85
Equivalence, 71
Esculin, 302
Esculetin, 302
Eserine, 298
Essence of mirbane, 293
Ether, 327
" acetic, 276
" ethyl, 275
" luminiferous, 29
" methyl acetic, 260

Hydrogen, arseniuretted, 150
“ basic, 77
“ chloride, 114
“ disodium phosphate, 184
“ fluoride, 111
“ light carburetted, 247
“ oxide, 97
“ peroxide, 110
“ phosphoretted, 146
“ sodium carbonate, 184
“ “ phosphate, 183
“ sulphate, 125
“ sulphuretted, 122
“ ammonium carbonate, 195
Hydrometer, 13
“ Twaddell, 13
Hydropotassium, carbonate, 191
“ oxalate, 192
“ sulphate, 192
“ tartrate, 193
Hydroquinone, 288
Hydrosodium sulphate, 183
Hydroxylamine, 136
Hygroscopic, 59
Hyoscyamine, 298
Hyponitrous oxide, 137
Hypothesis, definition of, 9
Hypoxanthine, 318
Hyposthenisants, 324

ICE, 99
Illuminating gas, 169
India rubber, 250
Indican, 302, 315
Imides, 291
Indigo, 315
“ carmine, 315
Indigotin, 315
Indium, 238
Induction, electrical, 37
“ coil, 46
Inebriants, 324, 327
Inertia, 14
Infectious diseases, 327
Infusion, 110
Inosite, 271
Invertin, 318–320
Iodine, 117, 326
“ compound tincture, 189
Iodine, oxides, 119
Iodoform, 261
Iridium, 240
Iron, 212, 326
“ cast, 213
“ chlorides, 216
“ chrome, 208
“ citrates, 217, 218
“ oxides, 214, 215
“ phosphates, 218
“ pig, 213
“ scale compounds of, 217
“ sulphates, 214
“ tartrates, 217
“ wrought. 213
Isomeric bodies, 246
Isomerism, 245

JALAPIN, 301
Jeweler's rouge, 215

KALIUM, 187
Kephalins, 317
Kermes' mineral, 160
Kerosene, 248
Ketones, 258
Kreatin, 317
Kreatinin, 317

LABARRAQUE'S solution, 186
Lac sulphuris, 122
Lactometer, 13
Lævogyrate, 268
Lamp black, 167
Lana philosophica, 206
Lanthopine, 298
Lapis infernalis, 236
Lardacein, 310
Laudanine, 298
Laudanosine, 298
Laudanum, 326
Laughing gas, 137
Law of Ampère, 17, 66
“ Avogadro, 17, 66
“ Berthelet, 84
“ Charles, 19
“ definite proportions, 92
“ Graham's, 21
“ Mariotte, 19
“ Ohm's, 44
Leaching, 110

Gasometer, 169
Gases, constitution of, 17
  " diffusion of, 20
  " poisonous, 327
Gastric juice, analysis of, 330
Gelatin, 311
Germicides, 132
Gersdorffite, 218
Giant powder, 266
Glass, 166
  " soluble or water, 166
Glauber's salt, 182
Globulin, 307
Globulins, 306
Glucoses, 267
Glucosides, 301
Glue, 311
Gluten, 310
Glycerine, 256
  " nitro, 266
Glycerines, 265
Glycerrhicin, 302
Glyceryl trinitrate, 266
Glycocin, 310
Glycogen, 272
Glycol, 256
Gmelin's test, 314
Gold, 237
  " trichloride, 238
Goulard's extract, 223
Graham's law of diffusion, 21
Gramme, definition, 15
Grape sugar, 270
Graphite, 166
Gravitation, definition of, 11
  " law of, 11
Gravities, specific, table of, 336
Gravity, specific, 11
  " " flask, 12
Greenockite, 207
Green, Brighton, 229
  " Brunswick, 228
  " mineral, 228
  " mountain, 228
  " Neuwieder, 228
  " Paris, 227
  " verditer, 228
Guaiacum, 252
Guanin, 318
Gum Arabic, 274
  " bassorin, 274
Gum, British, 272
  " lac, 253
  " resins, 252
  " Senegal, 274
  " tragacanth, 274
Gums, 252, 274
Gun cotton, 273
  " powder, 191
Gutta percha, 250
Gypsum, 199.

HÆMATIN, 308
  " hydrochloride, 308
Hæmin, crystals, 308
Hæmoglobins, 308
Hæmatite, 212, 215
Halogen elements, 111
Hausmanite, 210
Heat, 22
  " latent, 26
  " sensible, 28
  " specific, 27
Hematoïdin, 313
Hematoxylon, 316
Heavy spar, 202
Helleborein, 302
Helleborin, 302
Hemipeptone, 310
Hemlock, 327
Histozym, 318, 320
Homologous series, 245
Hornblende, 203
Hydrates, 77
Hydric chloride, 114
  " sulphate, 122
Hydrobilin, 314
Hydrobilirubin, 314
Hydrocarbons in air, 131
  " radicals, 255
  " homologous, nomenclature of, 247
  " homologous series of, 245
  " isologous series of 245
Hydrogen, 91
  " alcoholic, 258
  " and nitrogen, 134
  " and oxygen, 97
  " antimoniuretted, 158
  " arsenides, 150

Manganic compounds, 221
Manganite, 220
Manganous compounds, 220
Manna, 267
" Syrup of, 267
Mannite, 267
Mannitose, 267
Marble, 197
Mariotte, law of, 19
Marsh gas, 247
Mass, definition of, 10
Massicot, 220
Matter, definition of, 9
" divisions of, 10
" radiant, 18
" states of, 16
Measures, table of, 884
Meerschaum, 208
Melanin, 315
Melanogen, 315
Meconidine, 298
Melting point, 28
Menthol, 251
Mercurial tremors, 234
Mercuramidogen chloride, 230
" ammonium chloride, 232
Mercuric compounds, 231
Mercurous compounds, 230
Mercury, 229, 326
" ammoniated, 232
" fulminate of, 262
Metaldehyde, 257
Metalloids, 89
Metals, 89, 90
" alkaline, 180
" of the alkaline earths, 197
" platinum, 240
Methane, 247
" trichlor, 260
Methyl acetate, 260
" alcohol, 259
" chloride, 260
Metric system, 14
Micoderma cerevisiæ, 262
M crozimas, 320
Milk, composition of, 331
" of lime, 198
" sugar (lactose), 269
" of magnesia, 204
Mindererus, spirit of, 196
Mineral green, 228
Mineral waters, 309
Minium, 222
Mitis green, 228
Monamines, 265
Molecular attraction, 13
" weights, 66, 69
Molecules, 10, 63
" and atoms, multiplication of, 72
" composition of, 246
" compound, 76
" constitution, 246
" definition of, 63
" elemental, 67
" number of atoms in, 67
" salt, 77
" size and weight of, 17
Molybdenite, 241
Molybdenum, 241
Molybdic trioxide, 241
Monocalcium phosphate, 200
Monopotassium sulphate, 192
Monosodium phosphate, 185
Monsel's solution, 216
Morphine, 298, 326
Mortar, 198
" hydraulic, 198
Mountain green, 228
Mucin, 311
Mustard, 325
Mycoderma aceti, 279, 321
Myelins, 317
Myosin, 307
Myrosin, 318, 319
Myrrh, 253
Mythelene dichloride, 327

NAPHTHA, 248
Naphthaline, 253
Narceine, 298
Narcotics, 324, 326
Narcotine, 298
Nascent state, 93
Natrium, 181
Natural science, definition of, 9
Nessler's test, 136
Neurotics, 326
Neutrals, 76
Neuwieder green, 228
Niccolite, 218
Nickel, 218

Nickel compounds, 219
Nitre, 190
" spirits of, 141
" sweet spirits of, 276
Nitric oxide, 138
Nitriles, 293
Nitrites, 139
Nitro-benzol, 293, 327
Nitroform, 261
Nitro-glycerin, 266
Nitrogen, 128
" and hydrogen, 134
" chloride, 137
" dioxide, 138, 140
" iodide, 137
" monoxide, 137
" peroxide, 140
" protoxide, 137
" tetroxide, 140
Nitrous oxide, 137, 327
Nomenclature, 79
" examples of, 81
" irregularities in, 82
" of homologous hydrocarbons, 247.
" rule for, 79
" simplified, 80
Non-metallic elements, 90

OCCLUSION, 92
Ohm's law, 44
Oïdium albicans, 321
Oil of bergamot, 249
" of cloves, 249
" of copaiba, 249
" of cubebs, 249
" elemi, 249
" fusel, 264
" juniper, 249
" lavender, 249
" lemon, 249
" mineral sperm, 248
" orange peel, 249
" pepper, 249
" vitriol, 125
Oils, essential, 249
" fixed, 304
" lubricating, 248
Olefines, 248
Olein, 304
Oleomargarin, 305
Oleoresins, 252
Opium, 326
Ores, 205
Organic bodies containing nitrogen, 291
Organic compounds, synthesis of, 255
Organic chemistry, definition of, 243
Organic radicals, 254
Orpiment, 151
Osmium, 241
" tetroxide, 242
Osmosis, 57
Ossein, 311
Oxaluria, 201
Oxidation, 95
Oxygen, 94
Oxyhæmoglobins, 308
Oxymorphine, 298
Ozocerite, 248
Ozone, 96

PALLADIUM, 240
Palmitin, 304
Pancreatic juice, analysis of, 331
Papaverine, 298
Paradox, culinary, 24
Paraffin, 248
Paraffines, 247
Paraglobulin, 307
Paraldehyde, 257
Parchment paper, 126, 274
Paregoric, 326
Paris green, 152, 227
Pearlash, 191
Peat, 168
Pepsin, 319
Peptones, 310
Permanganates, 211
Pettenkofer's test, 312
Petroleum distillation, 248
Phenol, 287
Phloretin, 302
Phlorizin, 302
Phosphine, 146
Phosphorus, 143, 326
" black, 144
" oxides, 147
" oxychloride, 146
" pentoxide, 147

Phosphorus, red, 144
" trichloride, 146
" trioxide, 147
" physiological action, 145
" white, 144
Physical forces, 11
" science, definition of, 9
Physics, 9
" chemical, 16
Physostigmine, 298
Picnometer, 12
Picrotoxine, 299
Pilocarpine, 299
Piperine, 299
Plaster; -of-Paris, 199
" diachylon, 282
" lead, 282
Platinic tetrachloride, 241
Platinum, 240
" compounds, 241
" sponge, 240
Plumbago, 166
Plumbates, 221
Plumbic anhydride, 221
Plumboso-plumbic oxide, 221
Plummer's pills, 159
Poisons and their antidotes, 322, 325
" corrosive, 323
" irritant, 324
" neurotic, 324
" septic, 324
Poisoning, symptoms of, 322
" treatment of, 325
Poisonous gases, 324
Polarity, 37
" of the elements of batteries, 42
Polarization of plates, 42
Poles, 45
Polychroite, 302, 316
Polymerism, 245
Populin, 302
Potash, by alcohol, 188
" by lime, 188
" binoxalate of, 192
" neutral tartrate, 192
" red chromate, 209
" red prussiate, 179
" yellow prussiate, 179
Potassa, 188
" caustica, 188
Potassium, 187
" acetate, 192
" arsenite, 152
" bicarbonate, 191
" bromide, 188
" carbonate, 191
" antimonyl tartrate 160, 194
" chlorate, 189
" chloride, 188
" chromate, 209
" chrome alum, 210
" cyanide, 189, 327
" dichromate, 209
" ferricyanide, 179
" ferrocyanide, 179
" fluoride, 189
" haloid salts of, 188
" hydrate, 188
" hypochlorite, 190
" iodide, 189
" manganate, 212
" nitrate, 190
" oxalates, 192
" oxide, 187
" pentasulphide, 192
" perchlorate, 190
" permanganate, 212
" plumbate, 221
" quadroxalate, 192
" sulphate, 192
" sulphides, 192
" sulphites, 192
" tartrates, 192
Potato spirit, 264
Powder of algaroth, 159
Pressure, standard of, 20
Proteids, 305
Protagon, 317
Protopine, 298
Prussic acid, 176
Prussian blue, 179
Pseudomorphine, 298
Ptomaines, 300
Ptyalin, 318
Purple of Cassius, 238
Pyrolusite, 210, 211
Pyromorphite, 219
Pyroxylin, 273

QUANTIVALENCE; 71
" table of, 75
" variation in, 73
Quartz, 165
Quercitrin, 302
Quieting cordials, 326
Quinidine, 299
Quinine, 299
" sulphates, 299

RADICALS, 74
Radiometer, 18
Reactions, 83
" rules for writing, 85
" synthetical, 255
Reagent, definition of, 83
Realgar, 151
Red crocus, 215
Reduction, 93
Reinsch's test, 155
Rend rock, 266
Resin soap, 252
Resins, 252
" fossil, 252
" gum, 252
Resorcin, 288
Rheadine, 298
Rheophores, 45
Rhigolene, 248
Rhodium, 240
Rhodochrosite, 210
Rochelle salt, 193
Rubidium, 194
Ruby, 238, 239
Rum, 263
Rust, 213
Ruthenium, 240

SACCHARATES, 269
Saccharometer, 269
Saccharoses, 268
Saccharomycetes albicans, 321
Saffron, 316
Sal ammoniac, 194
" volatile, 196
Saleratus, 191
Salicin, 290, 299
Salicylates, 290
Saliva, composition of, 330
Salt, common, 182
" acid, 78
Salt, Epsom, 204
" Glauber's, 182
" of lemon, 192
" of Saturn, 223
" of tartar, 191
" Rochelle, 193
" Seidlitz, 204
" of sorrel, 192
Saltpetre, 190
" Chili, 185
Salts, 76, 77
" basic or sub-, 78
" double, 78
" metallic, 326
" normal, 78
Saponification, 265
Saponin, 302
Sapphire, 238, 239
Sarkin, 318
Saturated solution, 55
Scale compounds of iron, 217
Scales, thermometric, 28
Scheele's green, 152, 227
Schweinfurth green, 227
Science, definition of, 9
Selenium, 127
Septicine, 300
Serpentine, 203
Shell lac, 253
Siderite, 212
Signs, 74
Silicates, 166
Silicic hydride, 165
" oxide, 165
Silicium, 165
Silicon, 165
" compounds, 165
Silver, 235
" bromide, 236
" chloride, 236
" cyanide, 236
" fulminate, 262
" iodide, 236
" nitrate, 236, 326
" oxides, 235
" salts in photography, 236
Slag, 213
Smithsonite, 205
Soaps, 282
Soapstone, 203
Soda, 181, 183

Soda ash, 184
  " black ball, 184
  " caustic, 181
  " lye, 181
  " neutral carbonate, 183
  " sal. 183
  " washing, 183
Sodium, 181
  " benzoate, 289
  " borates, 164, 185
  " bromide, 182
  " carbonate, 183
  " chloride, 182
  " hydrate, 181
  " hypochlorite, 186
  " iodide, 182
  " nitrate, 185
  " oxides, 181
  " phosphates, 184, 185
  " potassium tartrate, 193
  " sulphates, 182
  " tungstate, 242
Solanine, 303
Solution, 55
  " copper, 53
  " Fehling's, 270
  " gold, 54
  " Lugol's, 118
  " Monsel's, 216
  " of gases, 55
  " silver, 54
Soot, 167
Sorbin, 267, 271
Sorbite, 267
Specific gravity, 11
Spectra, absorption, 35
Spectroscope, 32
Spectrum, 31
  " continuous, 33
  " solar, 33
Sphalerite, 205
Spirit of Mindererus, 196
  " methylated, 260
Spirits, 263
  " of nitre, 141
  " " sweet, 276
Sprue, 322
Starch, 271
  " paste, 272
Stearin, 304
Steel, 213
Steel, Bessemer, 213
Stercobilin, 314
Stibine, 158
Stings, 327
Stochiometry, 86
Storax, 253
Strontium, 201
Strychnine, 300
Sublimation, 25, 26
Substitution products, 254
Sucrates, 269
Sugar, barley, 269
  " cane, 268
  " grape, 270
  " inverted, 269
  " liquorice, 302
  " milk, 269
  " of lead, 223
Sugars, 268
Suine, 305
Sulphates, 126
Sulphocyanates, 178
Sulphur, 119
  " flowers of, 120
  " liver of, 192
  " roll, 120
Sulphuric ether, 275
Sulphurous chloride, 127
  " oxide, 124, 131
  " iodide, 127
Superphosphate of lime, 200
Sweat, analysis of, 332
Sylvite, 188
Synaptase, 319
Synthesis of organic compounds, 255
Synthetical reactions, 255
Syntonin, 307, 309
Syrup ipecacuanhæ, 325
  " soothing, 326

TABLE; of chemical elements, 64
Table; of compound radicals, 76
  " of electro-chemical series, 70
  " of measures, 335
  " of quantivalence of elements, 75
  " of solubility of chemicals, 337, 338

Table of specific gravities, 336
" of wave lengths, 31
" of weights and measures, 334
Talc, 203
Tannin, 303
Tartar, cream of, 193, 285
" crude, 193, 285
" emetic, 160, 194, 325
" soluble, 192
Taurine, 312
Teas, analysis of, 332
Tellurium, 127
Temperature, definition of, 27
" standard of, 20
Terra alba, 199
Test, Gmelin's, 314
" Marsh's, 156
" Pettenkofer's, 312
" Reinsch's, 155
Toxicological analysis, 157
Thebaïne, 298
Theobromine, 300
Theory, definition of, 9
Thermal unit, 27
Thermometers, 28
Thrush, 322
Tin, chloride, 326
Tincal, 185
Tobacco, 327
Tolu, 253
Toluene, 258
Topaz, 239
Torula, 262, 321
Treacle, 268
Tricalcium phosphate, 199
Tricupric carbonate, 227
Tungstates, 242
Tungsten, 241
Turmeric, 316
Turnbull's blue, 179
Turpeth mineral, 233
Turpentine, 249
Typical elements, 91
Tyrosine, 317

URINARY pigments, 314
Urine, analysis of, 332
Urine, black, 315
Urinometer, 13
Urobilin, 314
Uroxanthin, 314

VANILLIN, 301
Vapor, watery, 130
Venetian red, 215
Veratrine, 300
Verdigris, 228
Vermilion, 233
Vinegar, 279
Vitellin, 307
Vitriol, blue, 227
" green, 214, 326
" white, 206
Vulcanite, 250

WATER, 97
" alkaline, 109
Water, analysis of, 104
" chalybeate, 109
" glass, 166
" hardness of, 105
" ice, 102
" impurities of, 106
" laurel, 327
" lime, 198
" mineral, 109
" natural, 101
" of constitution, 204
" of crystallization, 59
" potable, 101
" purification of, 108
" rain, 101, 102
" saline, 109
" snow, 101
" spring, 102
" surface, 103
" thermal, 109
" vapor of, 101
" well, 102
Wax, animal, 248
Weight, atomic, 69
" molecular, 66, 86
Weights and measures, table of, 334
Whisky, 263
White lead, 223
" permanent, 223
" precipitate, 232
" vitriol, 206
Wine, 263
Witherite, 202
Wolfram, 241
Wood spirit, 278

Wounds, 327
Wulfenite, 219

XANTHIN, 317
Xanthophyll, 316
Xyloidin, 272

YEAST, 321
Yellow wash, 233

ZINC, 205, 207
amalgamation of, 42
Zinc, butter of, 206
" chloride, 206, 326
" hydrate, 206
" oxide, 206
" sulphate, 206, 325
" sulphocarbolate, 288
" toxicology, 207
" white, 223

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INDEX OF SUBJECTS.

	PAGE
CHEMISTRY AND PHYSICS	3
TECHNOLOGICAL AND MANUFACTURERS' BOOKS, ETC.	5
PHARMACY AND BOTANY	6
PHYSIOLOGY	9
THE MICROSCOPE	10
DYSPEPSIA, HEALTH, HAY FEVER, HEADACHES, ETC.	11
MISCELLANEOUS	12
HOSPITALS AND MANUALS OF NURSING	14
HYGIENE	14
THE CARE OF CHILDREN	16
HEALTH RESORTS	16

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